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1	85.48	91.67	96.01	92.22	92.78	93.06	86.02	92.62	93.61	93.01	88.89	91.40	348	1	196	1544
10	81.39	85.42	88.48	85.83	86.94	83.92	78.49	86.88	86.11	82.80	81.06	84.99	324	8	185	2043
100	70.76	79.79	78.68	76.39	77.50	69.84	68.61	83.33	74.72	72.85	71.72	74.93	323	13	178	2395
500	62.17	72.08	69.36	66.67	68.61	60.32	54.57	77.32	64.44	61.02	59.34	65.08	313	25	160	3542
1000	57.87	68.13	64.46	61.67	63.33	56.75	49.46	72.95	58.59	55.65	53.28	60.22	310	27	154	4137
2000	52.97	62.50	58.33	55.56	57.22	49.21	43.55	67.48	53.06	50.00	46.46	54.21	302	46	145	5005
3000	49.69	59.17	54.41	51.94	53.61	45.44	42.74	63.38	49.17	46.77	42.42	50.79	297	66	135	6180
4000	47.03	56.67	51.72	49.17	51.11	42.66	37.90	60.38	46.39	44.09	39.39	47.86	296	70	131	6750
5000	44.99	54.38	49.51	47.22	48.89	40.48	36.02	57.65	43.89	41.94	37.12	45.64	286	94	126	7516
6000	43.35	52.17	47.55	45.56	47.22	38.69	34.41	55.45	41.94	40.32	35.10	43.85	266	201	121	8494
7000	41.92	51.04	45.83	43.89	45.56	37.10	33.33	53.55	40.28	38.71	33.59	42.25	254	299	109	10848
8000	40.49	49.79	43.38	42.50	44.44	35.91	31.99	51.63	38.61	37.63	32.32	40.79	245	395	102	12874
9000	39.47	48.54	43.14	41.11	43.06	34.52	30.91	50.27	37.22	36.29	30.81	39.58	239	471	100	13215
10000	38.85	47.71	41.91	40.00	41.94	33.33	29.84	48.90	36.11	35.22	29.80	38.51	231	600	85	18547
QUAL%	38.85	47.71	34.55	33.88	35.00	23.81	20.16	22.50	25.00	25.00	17.42	29.44	212	1006	79	21157

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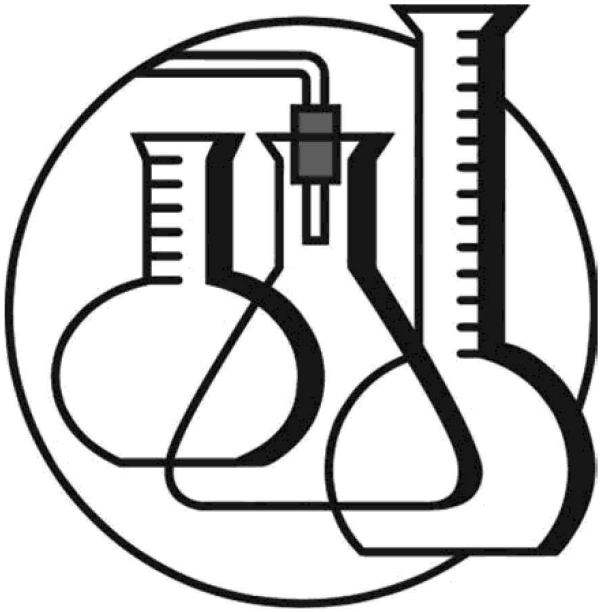
Chemical Bonding



INTRODUCTION

Nature is associated with innumerable chemical reactions. A chemical reaction involves change in molecular composition of a substance. It has been established that atom is the smallest particle of matter which takes part in a chemical reaction. Molecule is the smallest constituent particle of matter which has an independent existence and which represents the properties of the respective elements or compounds. Most of the elements in nature have been found to exist in combined state. However, there are some atoms which have independent existence and are considered to be highly stable atoms. Except those of a few elements, the atoms of most of the elements have an inherent tendency to combine and form molecules. The combining atoms may belong to the same element or different elements. Within the molecules, the atoms are held together by attractive forces. Study of the types of bonds, the conditions required for bond formation and the energy changes involved in chemical bond formation are some aspects of study of chemical bonding.

Reasons for the formation of a chemical bond

Noble gases are mono atomic. They do not combine either with their own atoms or atoms of other substances. These are the unique gases having 8 electrons (except helium) in their respective valence shells and is called the **octet configuration**. Helium has only two electrons as it has only one main energy level. This is called **duplet configuration**.

Except noble gases, atoms of all other elements have 1 to 7 electrons in their valence shells. The octet of the noble gases and the pair of electrons in helium are so stable that atoms of other elements also strive to attain the same electronic configuration to become stable.

In order to satisfy this urge, atoms unite to form molecules, their electrons get rearranged and a bond is said to be formed. During the formation of a bond, energy of the participating atoms gets reduced and thereby the molecules become stable.

Types of chemical bonds

The various types of bonds are:

1. Ionic bond
2. Covalent bond
3. Co-ordinate bond
4. Metallic bond

Ionic Bond

A chemical bond is formed between two atoms by complete transfer of one or more electrons from one atom to the other as a result of which the atoms attain their nearest inert gas configuration. The electrostatic force of attraction which holds the two oppositely charged ions together is called ionic bond.

During the formation of an ionic bond, one atom loses electron(s) forming positively charged ion called cation while the other atom gains electron(s) forming negatively charged ion called anion. The cation and anion are held together by strong electrostatic forces of attraction. The resulting compounds formed are called ionic compounds.

Example for ionic bond formation

Formation of potassium chloride

Electronic configuration of

Potassium (19) - 2, 8, 8, 1

Chlorine (17) - 2, 8, 7

Potassium has a tendency to lose its one valence electron to attain the nearest inert gas configuration of argon.

$K - e^- \rightarrow K^+$

Chlorine has a tendency to accept an electron from potassium to attain the electronic configuration of argon.

$Cl + e^- \rightarrow Cl^-$

$K^+ + Cl^- \rightarrow KCl$

Lewis dot formula

To explain the various types of bonds and to visualise the shift in the valence electrons, G. N. Lewis proposed the Lewis dot formula. In this, the valence electrons of the participating atoms are shown in the form of dot or cross. The valence electron of one of the participating atoms is represented as dot and that of the other one as cross.

Representation of some ionic compounds by Lewis dot formulae

Formation of magnesium oxide

Electronic configuration of

Magnesium (12) 2, 8, 2 

Oxygen (8) 2, 6. 

Magnesium loses two electrons from its valence shell while oxygen gains these two electrons such that both the elements attain the nearest inert gas configuration of neon.

Mg^{+2} 

Formula: MgO

Formation of calcium fluoride

Electronic configuration of

Calcium (20) → 2, 8, 8, 2 

Fluorine (9) → 2, 7 

Calcium loses two electrons from its valence shell while fluorine can gain only one electron to attain its nearest inert gas configuration. Hence the two electrons lost by one calcium atom are accepted by two fluorine atoms.

Ca^{+2} 

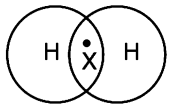
Formula: CaF_2

Covalent Bond

Covalent bond is a bond formed when two atoms share one or more electron pairs. Each atom contributes equal number of electron(s) towards the bond formation.

Formation of hydrogen molecule

Hydrogen has one electron. It shares this electron with another atom of hydrogen, attaining the duplet configuration. Hence the two hydrogen atoms share the electron pair and thereby a covalent bond is formed between the two hydrogen atoms.



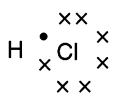
Formation of hydrogen chloride

Electronic configuration of

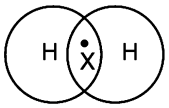
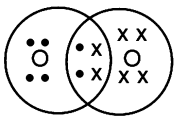
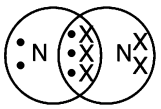
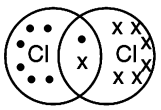
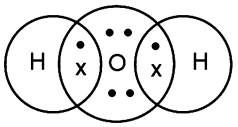
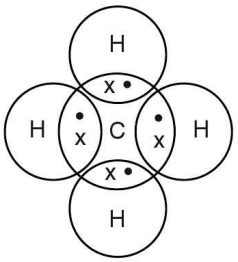
H (Z = 1) → 1 H •

Cl (Z = 17) → 2, 8, 7 

The pair of electrons (one is contributed by hydrogen and another one by chlorine) is shared by both hydrogen and chlorine atoms. Thus hydrogen attains its stable duplet configuration and chlorine attains its stable octet configuration.



Some examples of covalent molecules using Lewis dot symbols

Hydrogen(H_2)	**1**	 1 shared pair
Oxygen(O_2)	**6**	 2 shared pairs
Nitrogen(N_2)	**5**	 3 shared pairs
Chlorine(Cl_2)	**7**	 1 shared pair
Water(H_2O)	**H → 1, O → 6**	 2 shared pairs
Methane(CH_4)	**C → 4, H → 1**	 4 shared pairs

Factors responsible for the formation of ionic and covalent bond

Ionization energy	The lower the value of ionization potential of the element, the higher is the probability of cation formation. Likewise, higher ionization potential values lead to anion formation. Hence, if the difference of ionization potential between the two atoms is more, ionic compounds are formed.	Atoms with higher ionization potential are unable to lose their valence electrons and hence prefer to form covalent bonds by sharing of electrons.
Electron affinity	Atoms with very low electron affinity form ionic bond with the atoms of higher electron affinity.	The formation of a covalent bond is favoured when the combining atoms have almost equal electron affinity.

(Continued on following page)

Electronegativity	The greater is the difference in electronegativities between the two combining atoms, the greater are the chances of transfer of electron(s) from one atom to another. Hence, greater electronegativity difference between the two combining atoms leads to the formation of an ionic bond.	If the electronegativities of the combining atoms do not differ much, then the bond formed between them is likely to be covalent.
Metallic–Non-metallic character	If one of the atoms is metallic and the other one is non-metallic, then the difference in ionization potential, electron affinity and electronegativity becomes more which leads to the formation of an ionic bond.	If both the atoms are non-metallic, then the difference in ionization potential, electron affinity and electronegativity is very less, this leads to the formation of a covalent bond.

Energy changes during covalent bond formation

☛ *Example* Formation of hydrogen molecule.

When two hydrogen atoms approach each other, attractive forces develop between the electrons of one atom and the nucleus of the other atom. At the same time, repulsive forces also exist between the nuclei of the two atoms and the electrons of the two atoms. As the two atoms are brought closer to each other, at some distance, the proton electron attraction just balances the electron electron repulsion and proton proton repulsion. The bond formation takes place at this distance which corresponds to minimum possible energy state and the system becomes stable. This distance between the two nuclei is called the bond length.

The total energy of this system is a function of distance between hydrogen nuclei as shown in the graph.

In the hydrogen molecule, the electrons reside in the space between the two nuclei where they are attracted simultaneously by the protons present in both the nuclei of the two hydrogen atoms.

Since the bonding between two hydrogen atoms is associated with equitable sharing of electrons between the bonded atoms, this does not result in any charge separation within the molecule. These types of bonds are called **non-polar covalent bonds**.

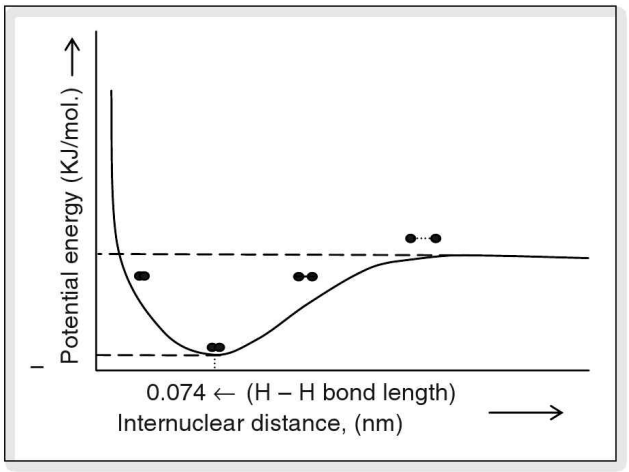


Figure 4.1

Polar Covalent Bond

Polar covalent bond is a type of covalent bond formed between two non-identical atoms. Since the two atoms differ in their capacity to attract the shared electron pair, unequal sharing of electrons results. Due

to the unequal sharing of electrons, fractional positive and negative charges are developed on the bonded atoms and the bond is said to be polar. Though it is a covalent bond, a slight ionic character is imparted to the bond due to the electronegativity difference between the two bonded atoms.

Comparative study between polar and non-polar covalent bonds

Sharing of electrons	Electron pairs contributed by both the bonded atoms of the molecule are unequally shared.	Electron pairs contributed by both the bonded atoms forming the covalent molecule are equally shared.
Development of charge	Slight positive and negative charges are developed on the bonded atoms	No charge is developed due to equal sharing of electrons.
Constituent atoms	Formed between the atoms having difference in their electronegativities. Example: H_2O, HF, HCl, NH_3.	Formed by identical atoms. Example: Cl_2, O_2, N_2 etc.

Comparative study of the properties of ionic and covalent compounds

Physical state	Most of the ionic compounds are crystalline solids. X-ray studies of the ionic compound have revealed that the constituent particles of the crystals are ions, not molecules. The cations and anions are held together very strongly by electrostatic force of attraction. Hence, the ions cannot be displaced from their position and thus the ionic compounds are generally hard solids.	Generally liquids or gases. This is because kinetic energy of the molecules easily overcomes the weak intermolecular (electrostatic) forces between the polar covalent molecules.	Generally liquid or gas. This is because kinetic energy of the molecules easily overcomes the weak van der Waal's forces acting between the non-polar covalent molecules.
Melting and boiling points	All the ionic compounds have high melting and boiling points. Considerable heat energy is required to overcome the electrostatic force of attraction between the ions of an electrovalent compound and make the ions mobile. Hence they have high melting and boiling points.	Have low melting and boiling points because a small amount of energy is sufficient to overcome the weak electrostatic force of attraction or the H-bonding between the polar covalent molecules.	Have low melting and boiling points because a small amount of energy is sufficient to overcome the weak van der Waal's forces acting between the molecules.

(Continued on following page)

Solubility	Ionic compounds are generally soluble in polar solvents like water, but insoluble in non-polar organic solvents. Molecules of the polar solvent like water can overcome the force of attraction between the ions in the crystal of the ionic solute. As a result, the ions become mobile and disperse in all directions in the polar solvent. Non-polar organic solvents cannot overcome the force of attraction between the ions of the ionic compound. Hence, ionic compounds do not dissolve in non-polar organic solvents.	Soluble in polar solvent due to the presence of partial charges. Also soluble in non-polar covalent liquids, due to similar forces between the molecules.	Insoluble in polar solvents like water because they don't ionize, but soluble in non-polar covalent liquids like benzene, carbon tetra-chloride, due to similar forces.
Density	The oppositely charged ions in an ionic compound are held closely by electrostatic force of attraction. Hence the number of ions per unit volume in an ionic compound is more and thereby their density is high.	Generally, they exist in the form of liquid or in gaseous states due to weak inter molecular forces. Hence the number of molecules per unit volume is less, thereby leading to low density.	Generally, they exist in liquid or gaseous states due to weak inter molecular forces. Hence the number of molecules per unit volume is less, thereby leading to low density.
Electrical conductivity	Electrovalent compounds conduct electricity either in the fused state or in their aqueous solutions, due to the presence of mobile ions. Since the ions are charged particles, they move towards the respective electrodes under the influence of an electric field and conduct electricity.	These compounds ionize in water and the ions help in conducting electricity.	These compounds do not ionize and hence do not conduct electricity.

Co-ordinate Covalent Bond

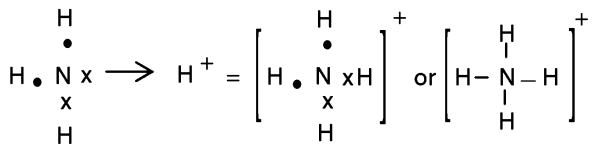
Co-ordinate bonds are those covalent bonds where the electron pair is provided by only one of the bonded atoms, but shared by both the atoms.

This atom which donates the electron pair is called the donor and the other atom which accommodates the shared pair of electrons is called the acceptor.

The co-ordinate bond is represented by a one sided arrow '→' where the arrow head points towards the acceptor and the tail towards the donor.

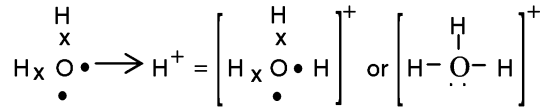
☛ *Examples*

(i) Formation of ammonium ion



The lone pair of electrons is donated by N-atom of ammonia and it is accepted by H^+. Hence N-atom is donor and H^+ is acceptor.

(ii) Formation of hydronium ion



O-atom of water molecule is donor and $H^+$is acceptor.

Forces acting between the covalent molecules

(i) van der Waal's forces

The weak forces which exist between the covalent molecules are known as van der Waal's forces.

There is an electrostatic force of attraction between the nucleus of one molecule and the electrons of the other. This is largely, but not completely neutralized by the electrostatic force of repulsion of electrons of one molecule by the electrons of the other or the nucleus of one molecule by the nucleus of the other. The resultant weak forces of attraction between the two molecules are called van der Waal's forces.

(ii) Dipole–dipole attractions

In polar covalent molecules, the unequal sharing of bonded electron pair results in partial charge separation within the molecule. The molecule with opposite partial charges is called dipole. The opposite charges of adjacent dipoles align with each other and the forces of attraction existing between these oppositely charged ends of the adjacent dipoles are called dipole–dipole attractions. The dipole–dipole attractions are much stronger than van der Waal's forces of attractions.

☛ *Example* In HCl molecule $H^{\delta+}$ and $Cl^{\delta-}$ charges are formed within molecule. In a sample of HCl, $Cl^{\delta-}$ of one HCl molecule attracts $H^{\delta+}$ of other molecule and these attractions are known as dipole – dipole attractions.

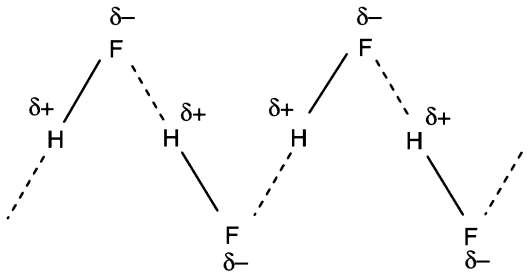
Hydrogen bonding

Polar covalent molecules which have a highly electronegative atom bonded to a hydrogen atom become strongly polar. Some examples of such molecules are HF, H_2O, NH_3.

In such types of polar covalent molecules, the H-atom of one molecule gets attracted to the strongly electronegative atom of the other molecule due to the formation of the slight positive charge on the hydrogen atom and the slight negative charge on the more electronegative atom.

This force of attraction that holds the hydrogen atom of one molecule to the highly electronegative atom of the other molecule is called hydrogen bond.

1.



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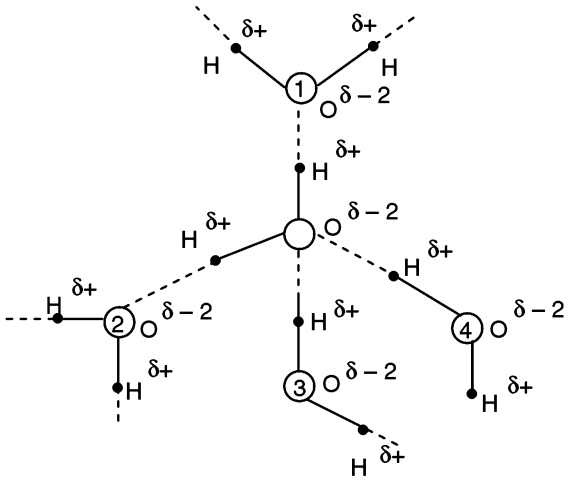

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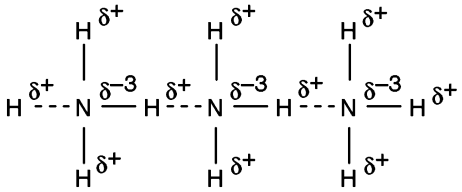


Metallic Bond

Any metal is a crystalline substance. A metal atom generally has either 1, 2 or 3 electrons in its valence shell. It can easily lose these electrons and gain stability. Metals are thus highly electropositive in nature. These electrons lost by the metal are called free electrons. The free electrons of all the metal atoms form an electron pool. The resulting positively charged metal ions are believed to be held together by the electron pool.

The force of attraction that exists between the mobile electrons and the metal ions is known as metallic bond.

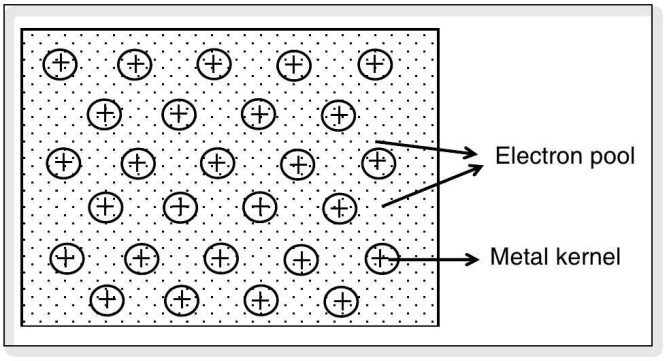


Figure 4.2 Schematic illustration of metallic bond in 'electron sea model'

Redox reactions

In a chemical reaction, if loss and gain of electron(s) takes place by the atoms or the ions, then the phenomenon can be defined in terms of loss and gain of electrons.

Oxidation

During a chemical reaction, if an atom or an ion loses one or more electrons, then the atom or the ion is said to be oxidized and this process is called oxidation.

Reduction

During a chemical reaction, if an atom or an ion gains one or more electrons, then the atom or the ion is said to be reduced and this process is called reduction.

Redox reaction

A chemical reaction in which loss and gain of electrons take place simultaneously is called a redox reaction.

Oxidizing agent: In the redox reaction, the atom or the ion that gains electron(s) is called an oxidizing agent.

Reducing agent: In the redox reaction, the atom or the ion that loses electron(s) is called a reducing agent.

Explanation of redox reactions with the help of some example

(i) When zinc granules are dipped into dilute H_2SO_4, H_2 gas is liberated with the formation of $ZnSO_4$.

$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$ (g)

In the above chemical reaction, the zinc atom loses two electrons and forms Zn^{+2} ion and each of the two hydrogen atoms gain an electron each.

$Zn - 2e^{-1} \rightarrow Zn^{+2}$

$2H^+ + 2e^{-1} \rightarrow H_2$

In this reaction, zinc is oxidized and acts as a reducing agent. H^+ is reduced and acts as an oxidizing agent.

(ii) When iron powder and sulphur are mixed and heated, they form ferrous sulphide (FeS).

$Fe + S \rightarrow FeS$

In this reaction, Fe loses two electrons and forms Fe^{+2} whereas S gets converted to S^{-2} by gaining two electrons. Hence, Fe is oxidized and S is reduced. Sulphur is an oxidizing agent whereas iron is a reducing agent.

test your concepts

Very short-answer type questions

1. How many types of chemical bonds are there? What are they?
2. Name one solvent in which most of the ionic compounds dissolve.
3. During the formation of _____ electron transfer takes place from one atom to the other.
4. In which state do the electrovalent compounds generally conduct electricity?
5. Why do the non-polar covalent compounds not conduct electricity?
6. What type of chemical bonds are found in each of the following compounds?
   (a) potassium chloride (b) carbon dioxide
   (c) hydrogen chloride (d) water
   (e) magnesium oxide (f) calcium fluoride
   (g) methane (h) sodium chloride
   (i) ammonia (j) phosphorus-penta chloride
   (k) sulphur hexafluoride
7. Ionic compounds are insoluble in _____ solvents.
8. In what type of solvents do the
   (a) polar compounds and
   (b) non-polar compounds dissolve?
9. The atoms of an element with electronic configuration 2, 8 are held by _____ forces of attraction.
10. Why are the molecules more stable than the atoms?
11. _____ compounds conduct electricity in fused state or in aqueous solution.
12. What are the criteria due to which a covalent bond becomes polar or non-polar?
13. What are the two opposing forces that start acting between two hydrogen atoms when they are brought together?
14. The nature of bond in H_2S is _____.
15. Why are the fractional positive and negative charge developed within a polar covalent molecule?
16. What is a coordinate covalent bond?
17. What are van der Waal's forces of attraction?
18. The electrical conductivity of ionic compounds in molten state is due to the presence of _____.
19. Define
   (a) polar covalent compound and
   (b) non-polar covalent compound

20. Give some examples in which coordinate covalent bond formation takes place.

21. Write two effects of van der Waal's forces of attraction.

22. Metallic lustre is due to the presence of _____.

23. Give some examples of the compounds where hydrogen-bond exists.

24. In the _____ bond, the contribution of an electron pair is one-sided, but the sharing is equitable.

25. Draw the Lewis dot formulae for the bond formations in the following compounds.

(a) NaCl (b) CaF_2 (c) H_2 (d) O_2

(e) N_2 (f) H_2O (g) NH_3 (i) CH_4

26. In a covalent bond, _____ of electrons takes place between the two atoms.

27. The chemical reaction in which loss and gain of electrons take place, is called _____ reaction.

28. Define the following terms on the basis of electron transfer.

(a) Oxidation
(b) Reduction
(c) Redox reaction
(d) Oxidizing agent
(e) Reducing agent

29. In ammonium ion, the lone pair of electrons is donated by _____ and it is accepted by _____. Hence _____ is called donor and _____ is called acceptor.

30. In the reaction $2Mg + CO_2 \rightarrow 2MgO + C$, oxidizing and reducing agents are respectively _____ and _____.

Short-answer type questions

31. Why is helium considered to be a noble gas though it has only two valence electrons unlike the other noble gases?

32. What are the factors responsible for the formation of covalent bond and ionic bond?

33. Why do the noble gases not take part in a chemical reaction?

34. What type of bond formation takes place between

(a) a metal and a non-metal and
(b) two non-metals?

35. Why do most of the ionic compounds exist in solid state while the covalent compounds are mostly in gaseous or liquid state?

36. What is the role of ionization potential and electron affinity of the elements in the formation of ionic bond and covalent bond?

37. Why is the density of the ionic compounds high and that of covalent compounds low?

38. Why are the melting points and boiling points of the ionic compounds high and those of covalent compounds low?

39. Why do the pure covalent compounds not conduct electricity?

40. Why are the ionic compounds soluble in water?

41. Why do polar covalent compounds dissolve in water?

42. Explain the hydrogen bond formation with the help of a diagram in the following compounds

(a) hydrogen fluoride (b) ammonia (c) water.

43. What types of bonds exist in the following ions?

(a) ammonium ion

(b) hydronium ion

Explain in brief their formation with a diagram.

44. Write in short about van der Waal's forces.

45. Mention the oxidizing agent, reducing agent in the following redox reactions and give reason in support of your answer.

(i) $Mg + 2HCl \rightarrow MgCl_2 + H_2$

(ii) $2ZnS + 3O_2 \rightarrow 2ZnO + 2SO_2$

(iii) $2KMnO_4 + 16HCl \rightarrow 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$

(iv) $2H_2S + SO_2 \rightarrow 2H_2O + 3S$

(v) $4HCl + O_2 \rightarrow 2H_2O + Cl_2$

Essay type questions

46. Write in detail about metallic bond.

47. Complete the following table:

Molecules	No. of valence electrons of the constituent atoms	Structure of the molecule (Present diagrammatically the sharing of electrons for covalent bond formation)
Chlorine		
Methane		
Water		
Nitrogen		

48. Differentiate between ionic and covalent bonding on the basis of the following parameters:

(a) formation
(b) physical state
(c) melting point and boiling point
(d) solubility
(e) electrical conductivity

49. Describe hydrogen bonding with the help of examples.

50. Describe the formation of co-ordinate bond and present diagrammatically the formation of coordinate bond in ammonium and hydronium ions.

CONCEPT APPLICATION

Concept Application Level—1

Direction for questions 1 to 7: State whether the following statements are true or false.

1. Atoms are less stable than molecules.

2. Sharing of electrons takes place in calcium fluoride.

3. In a nitrogen molecule, only one pair of electrons are shared between the two nitrogen atoms.

4. Metallic bond is unidirectional.

5. Covalent compounds conduct electricity since they have free electrons.

6. In a redox reaction, the oxidizing agent itself gets reduced.

7. In HF, unsymmetrical distribution of shared pair of electrons takes place.

Direction for questions 8 to 14: Fill in the blanks.

8. In calcium fluoride, _____ bond formation takes place between calcium and fluorine.

9. The total number of chemical bonds present in hydronium ion is _____.

10. The distance where the energy of the atoms taking part in the bond formation is minimum, is called the _____.

11. If the shared pair of electrons is present nearer to one of the bonded atoms, then the nature of the bond is _____.

12. Covalent compounds having giant molecules are virtually _____ in all solvents.

13. In water molecule _____ bond formation takes place between hydrogen and oxygen.

14. The nature of bonds formed between atoms of elements X and Y with higher ionization potential is _____.

Direction for question 15: Match the entries given in column A with appropriate ones in column B.

15.

A.	Ionic bond	( )	a.	Weak forces existing between molecules.
B.	Polar covalent bond	( )	b.	2 shared pairs of electrons
C.	Co-ordinate covalent bond	( )	c.	Electron transfer takes place from one atom to another
D.	Reducing agent	( )	d.	Attraction between positive ions and surrounding free mobile electrons
E.	Nitrogen	( )	e.	1 shared pair of electrons.
F.	Oxygen	( )	f.	Loses electron in a chemical reaction
G.	Chlorine	( )	g.	Shared pair of electrons is attracted towards the more electronegative atom
H.	Metallic bond	( )	h.	Forces of attraction between bonded hydrogen and a highly electronegative atom.
I.	H-bonding	( )	i.	3 shared pairs of electrons
J.	van der Waal's force	( )	j.	Contribution of pair of electrons by a single atom.

Direction for questions 16 to 45: For each of the questions, four choices have been provided. Select the correct alternative.

16. Among the following hydrides, ionic hydride is _____.

(1) MgH_2 (2) SiH_4 (3) BH_3 (4) PH_3

17. When one highly electropositive element A reacts with a highly electronegative element B, the compound formed will be

(1) an ionic compound.

(2) a polar covalent compound.

(3) a co-ordinate covalent compound.

(4) a non-polar covalent compound.

18. The covalency of nitrogen in ammonium ion is _____.

(1) 3 (2) 4 (3) 5 (4) 2

19. Which of the following substances is associated with the weakest electrostatic forces of attraction?

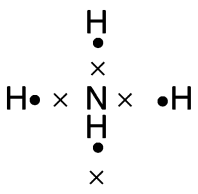
(1) HCl (2) NaCl (3) Na (4) Cl_2

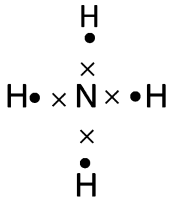
20. The force of attraction acting between cation and anion of an ionic compound is

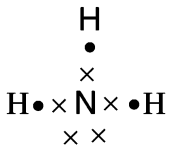
(1) electrostatic force of attraction (2) metallic bond

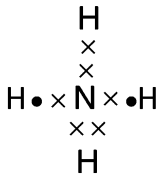
(3) hydrogen bond (4) None of these

21. The Lewis dot diagram representing ammonia molecule is

(1) 

(2) 

(3) 

(4) 

22. In the following reaction, $SO_2 + 2H_2S \rightarrow 3S + 2H_2O$

(1) sulphur is oxidized and hydrogen is reduced.

(2) hydrogen is oxidized and sulphur is reduced.

(3) sulphur is both oxidized and reduced.

(4) sulphur is reduced and oxygen is oxidized.

23. Though HCl, NH_3 are covalent molecules, their aqueous solutions conduct electricity, this is due to

(1) the presence of free electrons.

(2) the formation of free ions.

(3) the formation of hydrated compounds.

(4) formation of dative bond.

24. The strength of ionic bond is more when there is

(1) no difference in the atomic radii of the constituent atoms.

(2) less difference in the ionization potential of the atoms.

(3) less difference in the electronegativity of the atoms.

(4) greater difference in the metallic character of constituents atoms.

25. Why is the boiling point of ethyl alcohol (C_2H_5OH) higher than that of the corresponding hydrocarbon (C_2H_6)?

(1) Ionic bonds exist in ethyl alcohol molecule.

(2) Hydrogen bonds exist between ethyl alcohol molecules.

(3) Covalent bonds exist between ethyl alcohol molecules.

(4) None of the above

26. Which of the following solutions does not have a hydrogen bond?

(1) H_2S (2) C_2H_5OH (3) HF (4) NH_3

27. Which among the following is not attracted towards a charged plate?

(1) Water (2) Ammonia (3) Hydrochloric acid (4) Bromine

28. Assertion (A): Iron is harder than potassium.

Reason (R): Iron is more metallic than potassium due to its higher electropositivity.

(1) Both A and R are true and R is the correct explanation of A.

(2) Both A and R are true but R is not the correct explanation of A.

(3) A is true, R is false.

(4) A is false, R is true.

29. The nature of bonds present in sodium hydroxide are

(1) ionic, covalent and co-ordinate covalent.

(2) ionic and covalent.

(3) covalent and co-ordinate covalent.

(4) ionic and co-ordinate covalent.

30. If two atoms A and B of a molecule are brought closer than their minimum inter nuclear distance, then potential energy of the system

(1) remains constant at minimum value (2) starts increasing

(3) remains constant at maximum value (4) starts decreasing

31. Which among the following liquids shows convex meniscus in a glass tube?

(1) Water

(2) Hydrochloric acid

(3) Alcohol

(4) Carbon tetrachloride

32. The electronic configurations of two elements A and B are 2, 8, 8, 2 and 2, 6 respectively. Then the formula of the compound formed between them is _____ and its nature is _____.

(1) AB, covalent (2) AB, ionic (3) A_2B, covalent (4) A_2B, ionic

33. P, Q, R and S are four substances. P conducts electricity in the solid state, Q conducts electricity only in the solution state, R conducts electricity in the molten state and S is a bad conductor of electricity either in the molten state or in the solution state. Then P, Q, R and S may be _____.

(1) P = Aluminium, Q = $MgCl_2$, R = HCl, S = Br_2

(2) P = Aluminum, Q = HCl, R = Na_2O, S = Glucose

(3) P = KCl, Q = HCl, R = Iron, S = Glucose

(4) P = KCl, Q = HF, R = Na_2O, S = Br_2

34. Which among the following can form the strongest hydrogen bond?

(1) HF (2) H_2O (3) NH_3 (4) CH_4

35. A molecule in which the central atom is associated with contracted octet is

(1) NH_3 (2) PH_3 (3) $AlCl_3$ (4) CH_4

36. Covalency of N in NH_3, NH_4^+, NH_4Cl are ___, ___ and _____ respectively

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

37. Which among the following can form the strongest metallic bond?

(1) Sodium (2) Potassium

(3) Magnesium (4) Aluminium

38. Arrange the following in sequence for the representation of NH_3 molecule by Lewis dot diagram.

(a) Determination of the nature of bond between the constituents

(b) Electronic configuration of the constituents

(c) Representation of valence electrons as cross or dots

(d) Identification of the atomic numbers of the constituent elements

(1) dbca (2) dbac (3) bdac (4) dbca

39. "The strength of ionic bond in $MgCl_2$ is greater than in NaCl". Arrange the following key points that are essential to explain the above said statement in a correct sequence.

(a) Comparison of sizes of respective ions of Na and Mg.

(b) Factors affecting strength of ionic bond.

(c) Effect of size and electronegativity difference on the strength

(d) Comparison of electronegativity difference of the respective constituents in both the compounds

(1) d c b a (2) b c a d (3) b a d c (4) d a c b

40. Metals are lustrous in nature, having shiny appearance. Arrange the reasons given below in a sequence.

(a) Emission of radiation or light energy by excited electrons makes a metal shiny in appearance

(b) The electrostatic forces of attraction between metal ions and the mobile electrons is called metallic bond.

(c) The positive metal ions are surrounded by pool of electrons.

(d) When light falls on the crystal, electrons get excited

(1) c b a d (2) c b d a (3) b c a d (4) b d c a

41. The element which can never attain octet configuration in any of its compounds is _____

(1) K (2) Li (3) F (4) O

42. In the formation of AlF_3, aluminium atom has to lose _____ electrons.

(1) 1 (2) 2 (3) 3 (4) 4

43. Which of the following is a true statement?

(1) In polar compounds shared pair of electrons is away from the more electronegative atom.

(2) Polar compounds are good conductors of electricity in their vapour state.

(3) In polar compounds separation of charges take place.

(4) Polar compounds are good conductors of electricity in solid state.

44. In which of the following molecules partial charge separation does not take place?

(1) Chlorine (2) Hydrochloric acid

(3) Water (4) Ammonia

45. Ionic compounds do not conduct electricity in solid state. Identify the correct reason.

(1) Absence of oppositely charged ions in solid state

(2) Absence of mobile ions in solid state

(3) Absence of forces of attraction between ions in solid state

(4) Absence of free electrons in solid state

Concept Application Level—2

1. Explain the nature of the different types of bonds present in a NH_4Cl molecule.

2. The potential energy curve is not symmetrical about the minimum energy point. Justify.

3. Why is water called a universal solvent?

4. Comment on the intensity of charge of an electric field when HF and dry air are placed between two charged parallel plates.

5. Why are metals malleable and ductile?

6. Water shows capillary action in a narrow glass tube. Give appropriate reasons.

7. Between NaCl and CsCl, which has greater strength of ionic bond? Justify.

8. On decreasing the temperature, the conductivity of metals gradually increases. However below a particular temperature, the increase is found to be drastic. How do you account for this?

9. Which one of the two substances should have a higher boiling point–Br_2 or ICl? Give reasons in support of your answer.

10. The electrical conductivity of silicon increases by replacing a fraction of silicon atoms by arsenic atoms. Give appropriate reasons to support the statement.

11. Compound 'x' conducts electricity in the aqueous solution or molten state. Compound 'y' conducts electricity in the aqueous solution only. Compound 'z' does not conduct electricity in the molten state or in aqueous solution. Predict the nature of bonds in x, y and z.

12. Diamond, silicon carbide and silica can be used as abrasives. How do you account for the above said property?

13. "Though nitrogen and chlorine have almost equal electronegativity values, nitrogen forms hydrogen bonding, chlorine does not". Justify.

14. The leaves of aquatic plants do not decay though these are completely submerged. Explain.

15. What type of bond formation takes place in liquor ammonia? Explain.

Directions for questions 16 to 25: Application Based Questions

16. HF in the vapour state is associated covalent molecules while aqueous HF is ionic. Explain.

17. "Covalent bond is directional whereas ionic bond is not". Justify.

18. Compare NaCl and CsCl with respect to ease of formation and also the strength of the ionic bond.

19. Nitrogen and oxygen are the major components in air. However, they do not combine at all to form nitric oxide under normal conditions. Give reasons.

20. Intermolecular forces do not exist in ionic compounds. Justify the statement.

21. Compare and contrast the metallic character and hardness of sodium and iron. Justify.

22. Although the molecular mass of H_2S is more than that of H_2O, H_2O is a liquid whereas H_2S is a gas. Justify the statement.

23. Explain why when equal volumes of both ethyl alcohol and water are mixed, the volume of the resulting solution is less than the sum of the volumes of water and alcohol.

24. The leaves of aquatic plants do not decay though these are completely submerged in water. Explain

25. Identify redox reaction among the following

(1) Acid – base neutralisation (2) Precipitation reaction

(3) Metal displacement reaction (4) All the above

Concept Application Level—3

1. Why are ionic compounds hard and brittle?

2. Graphite is a good conductor of electricity whereas diamond cannot conduct electricity. Explain.

3. Density of water is maximum at 4°C. Explain.

4. Graphite is used as a solid lubricant. Why?

5. $SnCl_4$ is liquid while $SnCl_2$ is solid. Explain.

Directions for questions 6 to 10: Application Based Questions

6. Pure iron is relatively soft, ductile and malleable. But its hardness increases by diffusing carbon atoms in it. Explain with appropriate reasons.

7. Both zinc and mercury belonging to the same group in which differentiating electron enters into d-subshell of penultimate shell. But, zinc is a solid while mercury is a liquid. How do you justify this?

8. Three gases methane, ammonia and water vapour which have comparable molecular masses are liquefied at the same temperature. However, the pressure required to be applied is different for the three gases. Justify.

9. Among the various inorganic acids like HNO_3, $HClO_4$ and H_2SO_4, sulphuric acid is highly syrupy in nature. Explain.

10. Molten $AlCl_3$ is a poor conductor of electricity while hydrated $AlCl_3$ is a very good conductor. How do you account for this? Also explain the nature of bonding in the product.

Very short-answer type questions

1. Nature of forces of attraction

2. Any polar solvent

3. Ionic bond

5. Absence of charged species

6. Electro negativity, I.P., E.A. difference exist between constituent elements.

7. non-polar

10. Stability of lower energy state

11. Ionic

12. Charge separation

13. Attractive and repulsive forces

14. polar covalent bond

15. Difference in electronegativity

17. Weak forces existing between covalent molecules

18. mobile ions

20. NH_4^+ H_3O^+

21. (1) Boiling point (2) physical state

22. free electrons

23. H_2O, HF

24. co-ordinate covalent

25. Representation of
(i) valence electrons
(ii) sharing and transfer of electrons

26. sharing

27. Redox

28. (i) Loss of electrons
(ii) Gain of electrons
(iii) Mutual transfer of electrons
(iv) The substance that gains electrons
(v) The substance that loses electrons

Short-answer type questions

31. Only one shell (K shell).

32. Ionization potential and electron affinity

33. The presence of 8 electrons in valence shell.
(i) Ionization potential
(ii) Electron affinity and
(iii) Electronegativity

34. (a) complete transfer of electrons–ionic bond
(b) sharing of electrons–covalent bond.

35. Ionic compounds–strong electrostatic forces of attraction–solid.

Covalent compounds–weak electrostatic forces of attraction–liquid or gas.

36.

(a) I.P.	Large difference	High for both atoms
(b) E.A.	Large difference	Small difference

37. Strong electrostatic forces of attraction.

38. Ionic compounds–strong electrostatic forces of attraction.

39. Do not ionize.

40. They exist in the form of ions.

41. Charge separation.

42. (i) partial charge separation takes place on constituent atoms.
(ii) electrostatic force of attraction exist between H^{+s} and more electronegative element.

43. (i) Nitrogen donates its pair of electrons to H^+
(ii) Oxygen donates its pair of electrons to H^+

44. Weak electrostatic forces of attraction.

45. (i) loss of oxygen/gain of hydrogen–reduction
(ii) gain of oxygen/loss of hydrogen–oxidation.

Essay type questions

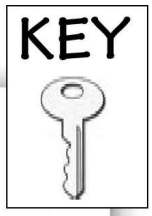
46. (i) Electropositive nature of metals
(ii) Free electrons
(iii) Electrostatic force of attraction

47. (i) Atomic number
(ii) Maximum number of electrons in a shell = $2n^2$

48. (i) Nature of bond
(ii) strength of bonds.

49. (i) Hydrogen bonded to electronegative atom
(ii) Charge separation
(iii) Electrostatic force of attraction

50. (i) Transfer of pair of electrons
(ii) Sharing
(iii) Donor, acceptor



Concept Application Level—1

True or false

1. True
2. False
3. False
4. False
5. False
6. True
7. True

Fill in the blanks

8. ionic
9. three
10. bond length
11. polar covalent
12. Insoluble
13. polar covalent bond
14. covalent

Match the following

15. A : c
B : g
C : j
D : f
E : i
F : b
G : e
H : d
I : h
J : a

Multiple choice questions

16. Choice (1)
17. Choice (1)
18. Choice (2)
19. Choice (4)
20. Choice (1)
21. Choice (3)
22. Choice (3)
23. Choice (2)
24. Choice (4)
25. Choice (2)
26. Choice (1)
27. Choice (4)
28. Choice (3)
29. Choice (2)
30. Choice (2)

31. Glass is made up of silicates which are ionic in nature. CCl_4 is covalent in nature, as a result, CCl_4 shows convex meniscus in a glass tube.

Choice (4)

32. Element B gains two electrons to attain octet configuration forming B^{-2}. Whereas element A loses two electrons to attain octet configuration and forms A^{+2}. Between A^{+2} and B^{-2}, there exists strong electrostatic forces of attraction (ionic bond) and forming AB.

Choice (2)

33. Generally, metals conduct electricity in solid state. Ionic compounds conduct electricity in the molten state as well as in the solution state. Polar covalent compounds conduct electricity in the solution state.
Non-polar covalent compounds conduct electricity neither in the molten state nor in the solution state and hence P = Al, Q = HCl, R = Na_2O, S = Glucose.

Choice (2)

34. The strength of hydrogen bond depends on size and electronegativity of an atom. The more the electronegativity and lesser the atomic size, the more is the strength of hydrogen bond. Hence HF can form the strongest hydrogen bond.

Choice (1)

35. In $AlCl_3$ central atom Al possess six electrons in its valence shell. Hence $AlCl_3$ is a molecule with contracted octet.

Choice (3)

36. Number of electron pairs involved in sharing will gives the covalency of an atom of an element. Hence covalency of 'N' in NH_3, NH_4^+ NH_4Cl are 3, 4, 4 respectively.

Choice (2)

37. The strength of metallic bond depends on number of valence electrons and size of metal kernel. The more the number of valence electrons and less the size of metal kernel, the more is the strength of metallic bond. Hence Al can form the strongest metallic bond.

Choice (4)

38. (i) Identification of atomic numbers of the constituents

(ii) Electronic configuration of the constituents

(iii) Determination of the nature of bond between the constituents

(iv) Representation of the valence electrons as crosses or dots

Choice (2)

39. (i) Factors affecting strength of ionic bond.

(ii) Effect of size and electronegativity difference on the strength.

(iii) Comparison of sizes of respective ions of Na and Mg.

(iv) Comparison of electronegativity difference of the constituents in both the compounds.

Choice (2)

40. (i) The positive metal ions are surrounded by a pool of electrons.

(ii) The electrostatic forces of attraction between metal ions and the mobile electrons is called metallic bond.

(iii) When light falls on the crystal, electrons get excited

(iv) Emission of radiation or light energy by excited electrons makes a metal shiny in appearance

Choice (2)

41. Li, as it attains stable electronic configuration of duplet in its compounds.

Choice (2)

42. Aluminium has to lose three electrons to form Al^{3+} ion.

Choice (3)

43. Due to electronegativity difference between the constituents in polar compounds separation of charges take place.

Choice (3)

44. In case of non-polar molecules separation of charge does not take place. Chlorine is non-polar.

Choice (1)

45. Though ionic compounds contain ions in solid state in their crystal lattices, they do not conduct electricity because the ions are at fixed positions and they are not mobile. Therefore, they do not act as charge carriers.

Choice (2)

Concept Application Level—2

Key points

1. Formation of ammonia molecule, ammonium ion and ammonium chloride.

2. (i) Intensity of force before bond formation.
(ii) Intensity of force at the minimum potential energy.
(iii) Intensity of force at where atoms move closer than bond length.

3. (i) Water has oxygen which is lightly electronegative.
(ii) Nature of solvent
(iii) Mechanism of dissolution

4. (i) Effect of field on polar molecules.
(ii) Orientation of polar molecules.
(iii) Effect of their orientation on electric field.

5. (i) Presence of free electrons.
(ii) Movement of free electrons
(iii) Directionality of metallic bond
(iv) Effect of directionality on malleability and ductility

6. (i) Type of molecules in glass and water.
(ii) Comparison of nature of bonds in water and glass
(iii) Forces acting between these molecules

7. (i) Factors affecting ionic bond strength.
(ii) Comparision of radius of cation
(iii) Effect of radius on electrostatic force of attraction

8. (i) Particles responsible for electrical conductivity in metals.
(ii) Factors that affect conductivity.
(iii) Relation between temperature and vibration of metal atoms.
(iv) Relation between vibration of metal atoms and conductivity.

9. (i) Inter molecular forces.
(ii) Comparing electronegativity of Br and Cl.
(iii) Nature of bonds in Br_2 and ICl based on electronegativity.
(iv) Effect of nature of bonds on intermolecular forces of attraction.

10. (i) Number of valence electrons in Arsenic.
(ii) Comparison of number of valence electrons in silicon and arsenic
(iii) Bonding in presence and absence of arsenic
(iv) Dependence of electrical conductivity on free electrons

11. (i) Co-ordinate compounds being covalent in nature, do not ionize.
(ii) Prediction of compound z
(iii) Comparing properties of x and y
(iv) Prediction of compounds x and y
(v) Predicting nature of bonds from type of compounds

12. (i) Nature of bonding in diamond, SiC and SiO_2.
(ii) Strength of the bond.
(iii) Requisite for abrasive nature.

13. (i) Sharing of electrons and transfer of electrons.
(ii) Comparison of atomic size of nitrogen and chlorine.
(iii) Effect of size on strength of hydrogen bond.

14. (i) Components and nature of outer layer of aquatic plants.
(ii) Bonding in water.
(iii) Forces between water and aquatic plants.

15. Formation of ammonium hydroxide.

16. Due to the greater difference in the electronegativity between 'F' and 'H' in HF the strength of H–bonding is strong even in the vapour state it exists in the associated form. (of 2 to 6 HF molecules).

But, HF in water exists in the form of ions because of high dielectric constant of H_2O, i.e. due to the high charge separation between O and H in H_2O, it breaks the polar covalent forces between H and F in HF and thus aqueous HF is an ionic compound.

17. Strong electrostatic forces of attraction exist between ions in all directions in the ionic bond. Hence it is non-directional in nature whereas electron pair is localized between the atoms in covalent bond which gives proper shape to the molecule and is directional in nature.

18. Ionic compounds are formed easily between larger cation and smaller anion. This is because an atom whose atomic radius is large can form a cation easily and an atom which has a small atomic radius can form an anion easily. Cs^+ ion can be formed more easily than Na^+ ion. Formation of CsCl is easier than the formation of NaCl. Among NaCl and CsCl, as the anion is the same i.e., Cl^-, the cations are compared. Cs has a larger atomic radius than sodium. Thus the strength of ionic bond is more in NaCl than CsCl.

19. N_2 is formed by sharing of three electron pairs, that is a triple covalent bond is present between two nitrogen atoms in N_2. The bond strength is very high since under normal conditions N_2 cannot react with any of the components present in air. Thus N_2 doesn't react with O_2 to form NO. High temperature favours the formation of NO.

20. Ionic compounds exist in the form of oppositely charged ions even in solid state. Therefore in ionic compounds, intermolecular forces do not exist, only interionic forces exist. These forces are so strong that the ions come together to form the crystal lattice.

21. Sodium belongs to IA group and has lower IP value than iron. Therefore, it can lose electrons more easily than iron. Hence it is more electropositive and more metallic than iron. Strength of the metallic bond increases with an increase in the number of valence electrons and effective nuclear charge. Thus iron is harder than sodium which contains only one valence electron and has less effective nuclear charges.

22. H_2O is a liquid while H_2S is not because oxygen being more electronegative than sulphur, can involve in the formation of hydrogen bonding in water. Thus water is a liquid while in H_2S there is no hydrogen bonding due to less electronegativity of sulphur and the molecules are far apart and hence H_2S is a gas.

23. The extent of intermolecular H–bonding between water and ethyl alcohol is greater compared to the extent of intermolecular H–bonding between ethyl alcohol molecules, i.e. aqueous or diluted ethyl alcohol has H–bonding to a greater extent compared to pure ethyl alcohol. Since each water molecule can form 4 H–bonds, many ethyl alcohol molecules get surrounded by it and thus more number of ethyl alcohol molecules are associated with water molecules thus decreasing the intermolecular spaces. Hence when equal volumes of both are mixed, the net volume is less than the sum of the two volumes taken. Thus, the ethyl alcohol molecules are brought closer in water due to greater extent of H–bonding.

24. The texture of aquatic plants (leaves, hydrocarbons stem, etc) is soft and waxy as they are made up of organic compounds, which are non-polar in nature. Since water molecules are highly polar, they hardly come in contact with the waxy surface of the leaves.

25. Only metal displacement reactions are redox reductions.

Choice (3)

Concept Application Level—3

1. (i) Forces of attraction in ionic compounds.
(ii) Type of constituents.
(iii) Arrangement of constituents.
(iv) Effect of this arrangement on directionality.

2. (i) Number of bonds formed by each carbon in graphite and diamond.
(ii) Number of valence electrons in carbon.
(iii) Comparison of the number of carbon atoms each carbon is bonded to in diamond and graphite.
(iv) Comparison of structure of diamond and graphite.

3. (i) Bonding in ice.
(ii) Change in bonding on melting.
(iii) Change in the position of molecules on increasing temperature up to 4°C.

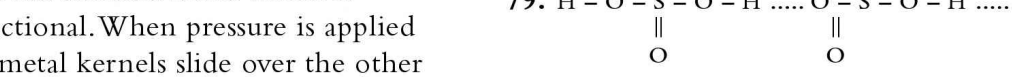
4. (i) Structure of graphite.
(ii) Bonding in graphite.
(iii) Bonding which helps in lubricative action.
(iv) Melting point of graphite.

5. (i) Comparison of charges on positive radial of $SnCl_2$ and $SnCl_4$.
(ii) Comparison of nuclear charge.
(iii) Effect of charge and nuclear charge on size.
(iv) Relation between size and intermolecular force of attraction.

6. Pure iron has metallic bond which is omnidirectional. When pressure is applied layers of metal kernels slide over the other layers and hence the metal can be made into thin sheets and wires. Therefore pure iron is malleable and ductile. When some amount of carbon is added a covalent compound Fe_3C (cementite) between iron and carbon is formed. Since covalent bonds are directional, layers of metal kernels can not slide easily. Hence the hardness of metal increases by the addition of carbon due to the formation of these bonds.

7. Zinc belongs to 4^{th} period and '3d' series of transition metals. The electronic configuration of zinc is 2, 8, 18, 2. It can contribute 2 electrons for metallic bond. Thus, strong metallic bond in zinc imparts hardness to the metal which makes its physical state to be solid. Mercury belongs to 6^{th} and '5d' series of transition metals. The electronic configuration of Hg is 2, 8, 18, 32, 18, 2. Due to the poor screening effect of 14 'f' electrons, the effective nuclear force of attraction on the outermost shell increases. Therefore, the electrons contributed for metallic bond are less which results in the formation of a thinner electron pool and hence weaker metallic bond. As a result of this, mercury is a liquid.

8. In H_2O and NH_3, there is hydrogen bonding and hence the intermolecular forces of attraction are stronger. Between NH_3 and H_2O, H_2O has stronger intermolecular forces of attraction than NH_3 due to stronger hydrogen bonding with oxygen. CH_4, being a non-polar molecule, has only van der waal's forces of attraction, which are weak. Therefore the order of critical temperatures is $CH_4 < NH_3 < H_2O$. The pressure required to be applied is in the order $H_2O < NH_3 < CH_4$. The greater the intermolecular force of attraction, the lower is the pressure required to liquify it.

79.


Since the electronegativity difference between S and O is more compared with N and O in HNO_3, Cl and O in $HClO_4$, the intensity of negative charge is more on oxygen and thus the hydrogen bonded to such oxygen can form hydrogen bonds more effectively. Thus H_2SO_4 molecules are associated with stronger H–bonding and form a syrupy liquid, i.e. a liquid with high viscosity.

10. $AlCl_3$ is covalent in nature because of small size and high nuclear charge of Al^{+3}.

But when $AlCl_3$ is put in water, Al^{+3} ions get surrounded by water molecules by means of coordinate covalent bonds in which a lot of energy is liberated.

$AlCl_3 + 6H_2O \rightarrow [Al(H_2O)_6]^{+3}$

Thus $AlCl_3$ gets dissolved in aqueous medium in which mobile hydrated aluminium and chloride ions are present and hence it conducts electricity due to the mobility of these ions in an electrical field. However, molten $AlCl_3$ exists as a molecule and thus is a bad conductor of electricity.