

بِسْمِ اللّٰهِ الرَّحْمٰنِ الرَّحِیْمِ
وَقُلْ رَبِّ زِدْنِيْ عِلْمًا



EDUCATION WITH HAMZA

**SUBJECT:
CHEMISTRY**

**CLASS:
11TH**

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Education with Hamza

وَقُلْ رَبِّ زِدْنِي عِلْمًا

DEDICATED TO

تمام تعریفیں رب لاشریک کے لئے جس نے مجھ ناپیز کو پیدا کیا۔ اور لاکھوں کروڑوں درود اس ذات پر جس نے ہماری بخشش کے لئے رورو کر دیا مانگی۔ ان نوٹس کو امت محمدیہ کی آسانی کے لئے بنایا گیا ہے۔ ان کا ثواب رب موسیٰ و ہارون نبی آخر زماں ﷺ کی جملہ آل و اولاد اور امت کو عطا کرے آمین۔

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Chemistry

Short Question

Chapter 2: Atomic Structure

Q1. What is the importance of Bohr's Atomic model? & Explain the charge & mass of fundamental sub-atomic particles?

Ans. (1) In 1913, Neil's Bohr's gave atomic model. This model describes electrons in first energy levels.

(2) He calculated the size of a hydrogen atom & the energy of its electrons in different shells.

(3) He also explained the spectrum of hydrogen. he said that when a electron jumps from higher to lower orbit, it emits energy in form of photon

	Mass/Kg	Charge (C)	Relative mass	Relative charge
Proton	1.672×10^{-27}	1.602×10^{-19}	1	+1
Neutron	1.675×10^{-27}	0	1	0
Electron	9.109×10^{-31}	1.602×10^{-19}	1/1836	-1

Q2. Explain the periodic trends of atomic radius with justifications? Why is the cation always smaller than the parent atoms & anion is bigger than the parent atom?

Ans. Periodic Trends

:In Periods: Atomic radii of ionic decreases from left to right along period.

Reason:- 1. It is due to increase in nuclear charge.

2. The shielding effect remains constant along period.

In groups: The atomic radii increases from top to bottom in a group.

Reason: 1. It is due to increase in the no. of shells. One shell is added in each group.

2. The anion are bigger in size than the parent atoms.

(i) Electron - electron repulsion increases due to increase in number of electron in the valence shell. Thus ionic size increases.

(ii) Due to increase in number of electrons decreases. Therefore, electrons move away & thus ionic size increases.

Q3. Why is the cation always smaller than the parents atoms. & ion is bigger than the parent atom? Explain how does different spectral series originate in hydrogen spectrum?

Ans. The cations are smaller in size than their parent atoms. It is due to reasons-

(i) Due to loss of electrons; effective nuclear charge on remaining electrons increases.

(ii) Due to loss of electrons the electron- electron repulsion decreases within valence shell

Emission spectrum of hydrogen-atoms-

- The emission spectrum of hydrogen is composed of distinct lines.
- These represent various electron transitions in the hydrogen atom b/w different levels.
- It helps to determine the energy levels of hydrogen & its electronic configuration.

Q4. Explain magnetic quantum number in detail. Why do s-orbitals have only one value of the magnetic quantum numbers? Why is it so that two electrons with same spin cannot reside in orbitals?

Ans. Magnetic Quantum number: The phenomena in which spectral lines split into multiple component in presence of a magnetic field is known as Zeeman effect. This effect can be explained based on the magnetic quantum number.

- It is also called "orientation quantum number".
- It is denoted by "m".
- Its possible values are from -l to +l .i.e means

$$m = \dots -3, -2, -1, 0, +1, +2, +3 \dots$$

If two electrons are present in same orbital (n, l, s, m are same) their spin will be opposite for both. If one is clockwise ($s = -1/2$) \uparrow blue others will be anti-clockwise ($s = +1/2$) \downarrow . e.g. two electron in 3p orbital have following sets of quantum numbers.

Electron A	Electron B
n= 3	n=3
l=1	l=1
m=-1	m=-1
s= -1/2	s=+1/2

Q5. Why 3d orbital has greater energy than 4s orbital? Explain (n+l) rule. Calculate the average atomic mass of magnesium keeping in view the relative abundance of its isotopes?

Ans. (n+l) Rules Here 'n' is principal quantum "l" is azimuthal quantum numbers.

- Higher the n+l value higher will be the energy of orbital.
- If two subshells have same (n+l) value then lower value of n lower will be the energy of subshell.

Example:

Subshell (n+l) value n-value Remarks

- | | | | |
|------|-------|---|-----------------------------|
| • 4s | 4+0=4 | 4 | Thus, 4s is of lower-energy |
| • 3d | 3+2=5 | 3 | |
| • 4p | 4+1=5 | 4 | Thus, 3d is of lower-energy |
| • 5d | 3+2=5 | 3 | |

From the spectrum following data can be obtained:

- Mg-24: Relative abundance = 79%, mass number = 24
- Mg-25: Relative abundance = 10%, mass number = 25

- Mg-26: Relative abundance = 11%, Mass number= 26.

So, Its average atomic mass will be :

$$\text{atomic mass (avg.)} = (24 \times 79) + (25 \times 10) + (26 \times 11) / 100 = 24.32 \text{amu}$$

Q6. What is effective nuclear charge? Describe the wavelength regions in which the spectral lines of H-spectrum fall.

Ans. The actual amount of positive nuclear charge experienced by an electron in a multi-electron atom is called effective nuclear charge.

$$Z_{\text{eff}} = Z - S$$

Where: Z= atomic number, S= shielding constant

According to bohr theory when hydrogen is heated or subjected to electric discharge, its electron moves from lower orbit 'n1' to higher orbit 'n2'. When this electron comes back, it emits some energy in form of photons of light.

Five series of spectral lines are present in hydrogen atom spectrum.

Q7. Explain ionization energy trends in the periodic table justification of these trends & anomalies & explore the practical application of atomic emission spectroscopy?

Ans. Trends in Groups: In a group, the ionization energy decreases from top to bottom.

Reasons: It is due to two factors: atomic size & shielding effect down the group atomic size increases. So, the valence electrons move farther away from the nucleus. The shielding effect also increases.

The Atomic Emission spectroscopy has been used to detect elements.

- Trace & major constituents in ceramics.
- Traces of Co, Ni, Mo & V in Graphite.
- Trace metal impurities in analytical reagents.
- Trace of Ca, Cu, Zn in blood.
- Zinc in pancreatic tissues.

Q8. Differentiate between atomic & ionic radius? & also its trends. & How does ionic radius varies from left to right in Periodic table?

Atomic	Ionic
The radius of an atom is defined as the average distance b/w the nucleus & its outermost electrons.	The radius of an ion is called ionic radius.
- Atomic radii & ionic radii decreases from left to right in periods. & it increases from top to bottom in a group	Ionic radii increase from top to bottom in a group due to increase in number of shell. For ions with same charge left to right within a period, the ionic radii decreases.

In a period from left to right, all elements does not form ions with same charge.

The charge on the ion depends upon electron configuration in the valence shell, which varies from left to right.

For ions with the same charge left to right within a period, the ionic radii decreases.

Q9. Define the following Principles with examples.

(a) Auf Bau Principles (b) Pauli Exclusion Principle- (c) Hund's rule

Ans. Pauli Exclusion Principle: "In an atom, no two electrons can possess identical values for all four quantum numbers. At least, their spin quantum numbers must differ."

Auf Bau Principles: Electrons are distributed within orbitals or subshells according to increasing energy levels. Initially, the lower-energy orbitals are filled then higher-energy orbitals are filled.

- Quantum mechanics allows us to calculate the energies of orbitals within an atom.
- The energy of orbital is decided by $n+l$ rule.

Hund's Rule: When degenerate orbitals are present, electrons prefer to occupy separate orbitals with the same spin rather than pairing up in the same orbital with opposite spin.

.Q10. Explanations of electronic configuration in term of electrons & inter electron repulsions?

Ans. The relationship between inter - electron repulsion & electronic contribution is: "Electron tend to minimize repulsion by occupying their own separate orbital rather than sharing an orbital with another electron."

Reasons:

- Electrons being negatively charged, repel each other due to their identical charges.
- To minimize this repulsion, electrons occupy separate orbitals within the same sub - shell.

Orbital energy: The energy of an orbital increases depends on its distance from the nucleus & the electron - electron repulsion it experiences.

Spin Pairing Energy: The energy of an orbital increases when two electrons with opposite spins occupy it.

Writing the electronic Configuration:

- The electronic configuration of elements involves filling electrons from lower energy levels to higher energy levels.
- The configurations $3p^6$ will mean $3p_x^2, 3p_y^2, 3p_z^2$, again the two electrons should be written as having clockwise \uparrow and anti-clockwise \downarrow spins.

Q11. Shapes of orbital?

Ans. Orbital is the region in space where the probability of finding an electron is maximum. Different orbitals have different shapes in three-dimensional space. p-orbital has two lobes, d-orbital has 4 lobes except d_{z^2} and S orbital is spherically symmetrical. The shapes of these orbitals..

- s-orbital: S orbital is a spherical
- p-orbital: P orbital is dumbbell shaped
- d-orbital: d orbital is cloverleaf shaped.
- f-orbitals: They have complicated shapes

The shapes of s, p and d orbitals are given below.

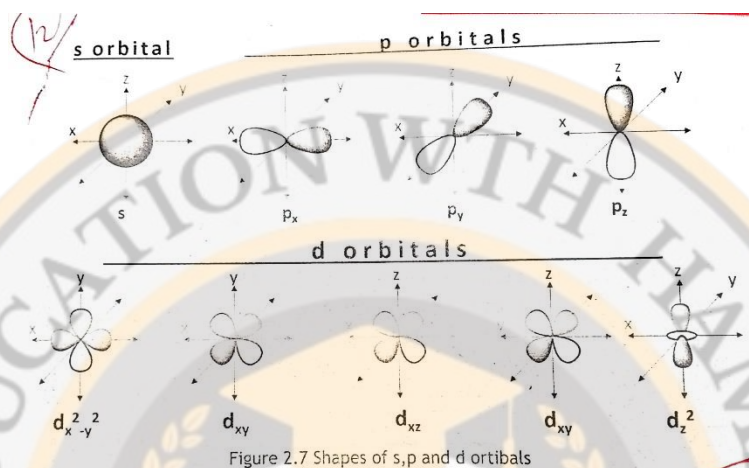


Figure 2.7 Shapes of s, p and d orbitals

Q12. Define Quantum number? Also define principal, Azimuthal and spin quantum number?

Ans. The digital/mathematical representation of the arrangement of electrons around the nucleus results in the development of quantum numbers.

Each electron within an atom is assigned a set of four numbers, known as quantum numbers

1. Principal Quantum Number

- This quantum number expresses the Bohr's shell in quantum form.
- It is represented by "n"
- It has possible values n=1,2,3...

Applications:

Period number:

The principal quantum number also gives the period number in the periodic table. Each period starts with the new shell. This number represents the Bohr's shell.

2. Azimuthal Quantum Number

- It describes the shapes of orbitals of electronic shells.
- It is denoted by "ℓ".

Orbital:

- An orbital is a space around the nucleus where the probability of finding the electron is maximum.
- One orbit or shell may consist of one or more than one orbitals or subshells.

3. Spin Quantum Number

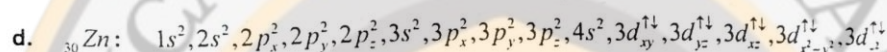
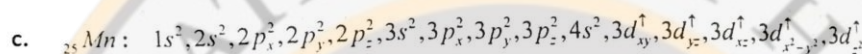
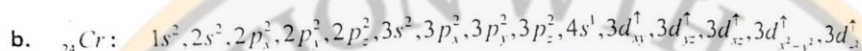
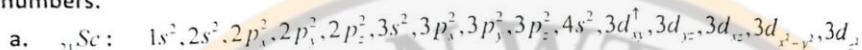
Spin of an electron is described by spin quantum number.

Maximum two electrons can live in one orbital formed by magnetic quantum number.

The two electrons occupy the same orbital, their spins must be opposite: if one electron spins clockwise, the other's spin will be counterclockwise.

Q13. Concept assessment.

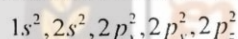
1. Write down the electronic configuration of the following elements representing all the four quantum numbers.



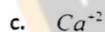
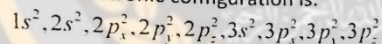
2. Write down the electronic configuration of the following ions.



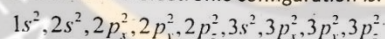
The atomic number of Na is 11. When Na^+ ion is formed then one electron is removed. So, Na^+ has 10 electrons. The electronic configuration is:



The atomic number of Cl is 17. When Cl^- ion is formed then one electron is gained. So, Cl^- has 18 electrons. The electronic configuration is:



The atomic number of Ca is 20. When Ca^{2+} ion is formed then two electrons are removed. So, Ca^{2+} has 18 electrons. The electronic configuration is:



NOTE: Both Cl^- and Ca^{2+} have same electronic configuration. So, they are isoelectronic species

Q14. Define mass spectrum? Write its principle?

Ans. Mass spectrometry (MS) is an analytical method employed to identify the chemical composition of a sample by measuring the mass-to-charge ratio of its ions.

Principle of Mass Spectrometry: Mass spectrometry involves three fundamental steps:

- ionization, mass analysis, and detection.
- The sample is first ionized, converting its molecules or atoms into charged ions.
- These ions are then separated based on their mass-to-charge ratio (m/z) and detected to generate a mass spectrum. The mass spectrum provides information about the composition and abundance of the sample's molecule or isotopic composition of element.

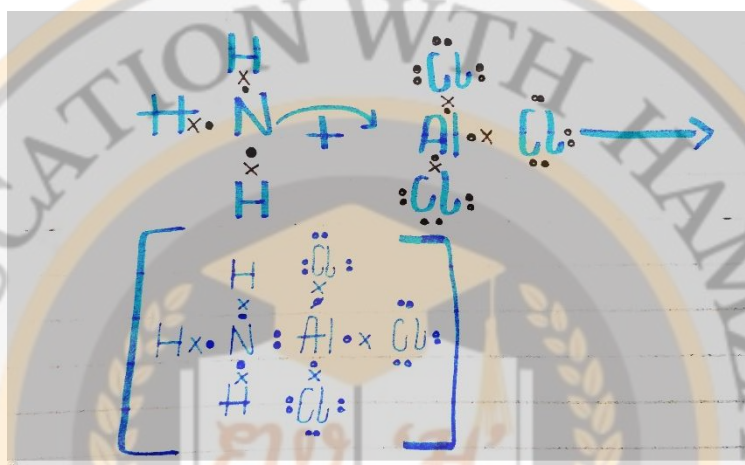
Chapter 3: Chemical Bonding

Q1. An ammonia molecule and an aluminum, chloride molecule can join together by the Forming a co-ordinate bond. (i) Explain how a co-ordinate bond is formed. Draw a dot & cross diagram to show the bonding in the compound formed between ammonia & Aluminum chloride, H_3NAICl_3 .

Ans.

- Explain how a co-ordinate bond is formed.
- NH_3 has three covalent bonds & one electron pair on nitrogen atom.
- In $AlCl_3$ octet of Al is not complete. It has six electrons.

(ii) Draw a dot & cross diagram to show the bonding in the compound formed between ammonia & Aluminium chloride, H_3NAICl_3



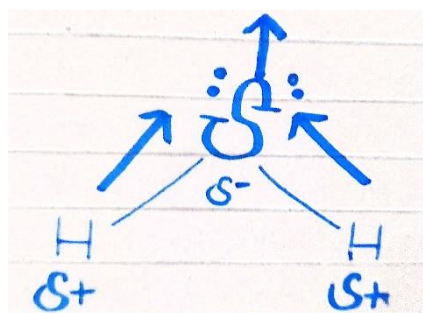
Q2. Hydrogen sulphide(H_2S) is a covalent compound. Explain the type of hybridization also write bond angle is HSH .Also show on your diagram the partial charges on each atom as $\delta+$ on H & $\delta-$ on S an arrow showing the exact direction of dipole in molecule as a whole? Oxygen 'O', sulphur S, selenium Se, are in the same group in the periodic Table.

- (i) explain why hydrogen selenide, H_2Se , has a higher boiling than hydrogen sulphide, H_2S .
- (ii) Explain why the boiling point of water is so much higher than the boiling of hydrogen sulphide.

Ans. In H_2S , the sulphur atom may be sp^3 -hybridized.

Four sp^3 hybrid orbitals are presents at the corner of regular tetrahedron.

Two half-filled sp^3 -hybrid orbitals of sulphur overlap with half-filled 1s orbitals of two H atoms to form two sigma bonds.



(i) Explain why hydrogen selenide....

The electronic cloud size increases down the group in periodic table due to increase in atomize / number.

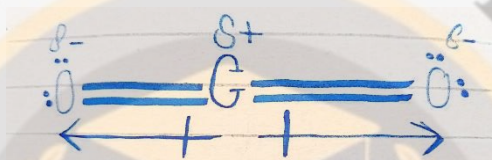
(ii) Explain why the boiling point of water:

In H_2O strong hydrogen bonding is present two hydrogen bond per molecule are present in water. In H_2S only weak dipole-dipole forces are present.

Q3. Describe the shape of the carbon dioxide molecule:-

Ans. In CO_2 the carbon atom is sp -hybridized.

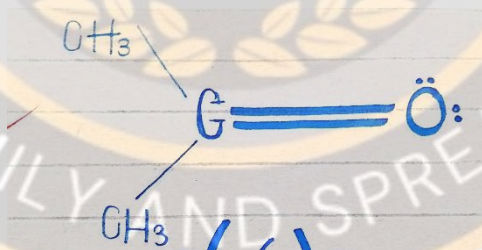
It has two linear sp -hybrid orbitals which overlap with sp^2 orbitals of two oxygen atoms. Further, two un-hybridized p -orbitals of carbon atoms overlap sideways with p -orbitals of both oxygen atoms to form π -bonds.



Q4. Explain why ice is less dense than liquid.

Ans. Most of the solids are denser than their liquids. This is because in solid states, the molecules are more closely packed. But in ice, there is a three-dimensional hydrogen bonded network of water molecules. The hydrogen bonds are long so water molecules are farther apart from each other in ice. This produces a rigid lattice but more 'open' arrangement with empty spaces

Propanone has the structure shown below-



Q5. CO_2 & CS_2 have linear structure. Justify. & NH_3 & BF_3 are both tetra-atoms molecules. The dipole moment of BF_3 is zero while that of NH_3 is 1.49D.

Ans. In CO_2 & CS_2 , there are four electron pairs around the carbon atom. In a double bond, the electron pairs must be present between the two atoms just like electrons of single bond. They cannot go to different place in valence shell.

NH_3 has dipole moments, which shows that the individual bond moments do not cancel the effect of each other. It is because NH_3 has pyramidal structure. BF_3 has zero dipole moments which shows that the individual bond moments cancel the effect of each other. It has the triangular planar structure.

Q6. The bond orders of O₂ is less than O₂⁺. CO₂ - molecules have linear structure while He₂ have angular structure. Comment:

Ans. In O₂ there are 6 electrons in valence bonding molecular orbitals & two electrons in valence ant bonding molecular orbitals. Thus

$$\text{Bond Order} = \frac{6-2}{2} = 2$$

In O₂⁺ one electron is lost from ant bonding molecular orbital. Therefore there are 6 electrons in valence bonding molecular orbitals & only one electrons in valence ant bonding molecular orbital. So, its bond order is.

$$\text{Bond Order} = \frac{6-1}{2} = 2.5$$

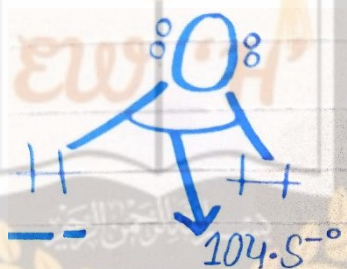
Hence,

bond order of O₂⁺ is more than O₂.

The O₃ molecule has no lone pair.

According to VSEPR theory, both double bonds are present at the angle of 180° in order to have minimum repulsion. Hence,

It has linear structure.



Q7. Differentiate between sigma & Pi bond. & Also differentiate between Atomic & molecular orbitals.

Sigma-bond	Pi-bond
It is formed when two atomic orbitals overlap with their axes in a straight.	It is formed when two p-orbitals overlap with their axes parallel to each other.
The overlap of atomic orbitals is more in sigma bond. Hence, a big amount of energy is released & bond is strong.	The overlap of atomic orbitals is very slightly. Hence, this bond is quite weak.

Q8. Distinguish b/w diamagnetic and paramagnetic substances?

Diamagnetic substances	Paramagnetic substances
These are weakly repelled by strong magnetic field.	These are weakly attracted by strong magnetic field.
These do not have unpaired electrons.	These have unpaired electrons
The paired electrons cancel the magnetic moments of each other. So, these have zero magnetic moment.	. The magnetic moment of unpaired electrons is not cancelled thus they have magnetic moments. Increase in number of unpaired electrons increases the magnetic moment
e.g. Iron (Fe)	e.g. Zinc (Zn)

Q9. Why MOT is superior to VBT?

Valence Bond Theory	Molecular Orbital Theory
It does not explain the paramagnetic behaviour of molecules e.g., O ₂ molecule	It explains the paramagnetic behaviour
It does not give idea about the bond order	It gives idea about bond order. So we can determine whether the bond is single, double or triple or even in fraction
It does not explain the non-existence of He ₂ molecule	It explains the non-existence of He ₂ molecule

Chapter-4: Stoichiometry

Q1. 49g each of H₂SO₄ & H₃PO₄ have same numbers of molecules but having different numbers of atoms. & Different gases having different masses occupy equal volume at S.T.P. Why?

Ans. Number of moles of H₂SO₄ = 49/98 = 0.5 moles

Number of moles of H₃PO₄ = 49/98 = 0.5 moles

Since numbers of moles are same for both H₂SO₄ & H₃PO₄, therefore, they contain the same number of molecules.

In gases distance between two molecules is much greater than their molecules size. Thus volume occupied by gas molecules does not depend upon the size or mass of molecules. It only depends upon the numbers of molecules. Hence, different gases having different masses occupy equal volume at STP.

e.g. 22.414 dm³ of H₂ gas at STP = 1 mole = 2g = 6.02 × 10²³

22.414 dm³ of NH₃ gas at STP = 1 mole = 17g = 6.02 × 10²³

Q2. One mole of Na & Al has equal number of atoms but size of atoms are different. why? b limiting reactant is always in lesser quantity mixture or not is why?

Ans. One mole of a substance always contains Avogadro's number of molecules. So, one mole of both Na & Al contains 6.02 × 10²³ molecules.

The reactant which consumes first in a chemical reaction is called limiting reactant.

(i) Number of moles of each reactant is calculated from given masses.

(ii) Using balanced chemical equations amount of product, expected to form, from each reactant is calculated.

Q3. Amount of product obtained through balanced chemical equation is greater than the amounts obtained experimentally? Sp What are the basic assumptions in stoichiometric calculations?

Ans. It is due to following reasons.

(1) Formation of by-products due to side reaction.

(2) Reversible reactions may not go to completion.

(3) Mechanical loss due to filtrations, distillations & separations etc.

- (1) Reactants are completely converted into product.
- (2) No side reaction occurs.
- (3) Law of conservation of mass & law of definite proportions is obeyed

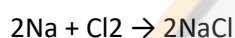
Q4. 18g of steam has Avogadro's numbers of molecules but 58.5g NaCl has not. & why 2 moles of Na react with 1mole of chlorine gas to produce 2 moles of NaCl?

Ans. Steam (H₂O) is a molecular substance in which H₂O molecules exist independently.

So, 18g H₂O = 1 mole = 6.02×10^{23} molecules of H₂O

NaCl, is an ionic substance which consist of Na⁺ & Cl⁻ ions. Since, NaCl cannot exist as independent molecules therefore term formula unit is used to express numbers of particles of NaCl.

This is according to the balanced chemical equations of formation of NaCl.



Q5. Calculate each of following quantities.

Ans. (1) Mass in gram of 0.74 mol KMnO₄.

Number of moles of KMnO₄ = 0.74 mol

Molar mass of KMnO₄ = 39+55+16x4 = 158g_{mol}⁻¹

Mass of KMnO₄ = No. of moles × Molar mass

$$= 0.74 \times 158 = 116.92\text{g}$$

(2) Mass of atoms in 9.22g Mg(NO₃)₂

Given mass of Mg(NO₃)₂ = 9.22g

Molar mass of Mg(NO₃)₂ = 24+14x2+16x6 = 148g_{mol}⁻¹

Number of moles of Mg(NO₃)₂ = Given mass / Molar mass

$$= 9.22 / 148 = 0.062 \text{ moles.}$$

(3) Mass in kg of 2.6×10^{20} molecules of SO₃.

Number of molecules of SO₃ = 2.6×10^{20}

Molar Mass of SO₃ = 32+16x 3 =80g_{mol}⁻¹

number of molecules of SO₃ = Mass / Molar mass x NA

$$2.6 \times 10^{20} = \text{Mass} \times 6.02 \times 10^{23} / 80$$

$$\text{Mass} = 2.6 \times 10^{20} \times 80 / 6.02 \times 10^{23} = 3.46 \times 10^{-2} \text{ g}$$

$$\text{Mass in kg} = 3.46 \times 10^{-2} / 1000 = 3.46 \times 10^{-5} \text{ kg}$$

(4) Mass in mg of 0.48 mole of

CuCl₂.2H₂O

Number of moles of CuCl₂.2H₂O= 0.48mol

Molar mass of $\text{CuCl}_2 \cdot 2\text{H}_2\text{O} = 63.5 + 35.5 \times 2 + 1 \times 4 + 16 \times$

$2 = 170.5 \text{ g mol}^{-1}$

So,

Mass of $\text{CuCl}_2 \cdot 2\text{H}_2\text{O} = \text{Number of moles} \times \text{Molar mass}$

$= 0.48 \times 170.5 = 76.73 \text{ g}$

mass of $\text{CuCl}_2 \cdot 2\text{H}_2\text{O} = 76.73 \times 1000 = 76730 \text{ mg}$

(5) Volume of SO_2 at STP of 4.8×10^{23}

molecules of SO_2 .

Number of molecules of $\text{SO}_2 = 4.8 \times 10^{23}$

Since:

1 mole of $\text{SO}_2 = 6.02 \times 10^{23}$ molecules of $\text{SO}_2 = 22.4 \text{ dm}^3$ at STP

So, 6.02×10^{23} molecules of $\text{SO}_2 = 22.4 \text{ dm}^3$ at

STP

4.8×10^{23} molecules of $\text{SO}_2 = 4.8 \times 10^{23} \times 22.4 /$

$6.02 \times 10^{23} = 17.87 \text{ dm}^3$ at STP.

(6) Covalent bond in 22 gram of dry ice-

Dry ice is CO_2 .

Given mass of $\text{CO}_2 = 22 \text{ g}$

Molar mass of $\text{CO}_2 = 44 \text{ g mol}^{-1}$

Number of molecules of $\text{CO}_2 = \text{Mass} \times N_A / \text{molar mass}$

$= 22 \times 6.02 \times 10^{23} / 44 = 3.01 \times 10^{23}$ molecules

The structural formula of dry ice (CO_2) is

$\text{O} = \text{C} = \text{O}$

Thus,

1 molecules of CO_2 contains = 4 covalent bonds

3.01×10^{23} molecules of CO_2 contains = $4 \times 3.01 \times 10^{23}$

$= 1.204 \times 10^{24}$ covalent bonds

Q6. How many covalent bonds are present in 9g of H_2O ? Differentiate between limiting & non-limiting?

Given mass of water = 9g

Molar mass of water = 18 g mol^{-1}

Therefore,

Number of molecules of water = Mass of water \times N_A / Molar mass of water

Number of molecules of water = $9 \times 6.02 \times 10^{23} / 18 = 3.01 \times 10^{23}$

Thus 1 molecule of H₂O contains covalent bonds = 2

3.01×10^{23} molecules of H₂O contain covalent bonds =

$2 \times 3.01 \times 10^{23} = 6.02 \times 10^{23}$ bonds.

Limiting	Non-limiting
The reactant that is consumed completely in a chemical reaction is called "limiting reactant".	The reactant left un-reacted after the completion of reaction is called "non-limiting".
The limiting reactant controls the amount of product formed in a chemical reaction.	The non-limiting reactant does not control the amount of product.
The amount of product formed depends upon of initial amount of limiting reactant.	The amount of product formed is independent of initial amount of non-limiting reactant.

Q7. Mg atom is twice heavier than that of carbon atom. Justify it. Explain why balanced chemical equations are used in stoichiometric problems?

Ans. One atom of carbon contains 6 protons & 6 neutrons in its nucleus & its atomic mass on atomic mass unit scale is 12 amu.

While one atom of Mg contains 12 protons & 12 neutrons in its nucleus & its atomic mass on atomic mass unit scale is 24 amu.

Thus,

Mass of Mg = 24 amu = 2

Mass of C = 12 amu

or Mass of Mg = 2x Mass of C

Hence,

One atom of Mg is twice heavier than that of one atom of carbon.

Stoichiometric calculations are done with the balanced chemical equations. In this assumed that matter is neither created nor destroyed during a chemical change.

Hence, law of conservation of mass has to be considered during stoichiometric calculation.

Q8. Difference between Actual Yield and Theoretical Yield? Also define Molar volume?

Actual Yield	Theoretical Yield
It is the amount of a product actually obtained during a chemical reaction.	It is amount of a product calculated from a balanced chemical equation.
It is always less than the theoretical yield.	It is always greater than actual yield.
Reversible nature of the reaction & side reactions decreases then actual yield.	while calculating theoretical yield. 1. No side reactions occurs. 2. Reactants are completely converted into products Therefore, yield is not decreased.

Molar Volume (V_m):

One mole of any gas at STP occupied a volume of 22.414 dm^3 . This volume is called "Molar volume."

22.414 dm^3 of any gas at STP 1 mole = 6.02×10^{23} molecules

22.414 dm^3 of H_2 gas at STP = $2\text{g} = 6.02 \times 10^{23}$ molecules

22.414 dm^3 of NH_3 gas at STP = $17\text{g} = 6.02 \times 10^{23}$ molecules.

Q9. Define solution stoichiometry & importance of stoichiometry?

Ans. In solutions the most common concentration unit is molarity. The molarity of a solution is simply the numerical value of its concentration in mol/dm^3 .

"Molarity is defined as numbers of moles of solute dissolved per dm^3 of solution."

Mathematically,

$M = \text{mole of substance} / \text{dm}^3 \text{ of solution}$

Importance-

- In the preparations of antibiotics, the stoichiometry ensures that each does matches the active ingredient & target bacteria.
- Use of insulin relies on the stoichiometry to precise control of blood sugar levels.
- Stoichiometry determines the concentration of viral antigens in preparation of vaccine for affective result.

Q10. Determine the volume that 2.5 moles of chlorine molecules occupy at STP.

Ans. Solution:

Since

1 mole of Cl_2 occupies a volume of 22.414 dm^3 at S.T.P., therefore

1 mole of Cl_2 gas occupy a volume = 22.414 dm^3 of Cl_2 at S.T.P.

2.5 mole of Cl_2 occupy a volume = $2.5 \times 22.414 \text{ dm}^3$

= 56.035 dm^3

Q11. Calculate the gram molecular mass of a gas which has density of 1.43 g/dm^3 at STP.

Ans. Solution:

Density of gas is 1.43 g/dm^3 . It means 1 dm^3 of gas contains 1.43 g

So

1 dm^3 of gas at STP = 1.43g

22.414 dm^3 of gas at STP = 1.43×22.414

= 32.052g

Since

22.414 dm^3 of a gas at STP = 1 mole of gas = Molar mass of gas

Therefore, gram molecular mass of gas is 32.052amu.

Q12. How many moles of oxygen molecule are there in 20.0 dm³ of oxygen gas at STP?

Ans. 22.414 dm³ of oxygen gas at STP = 1 mole

1 dm³ of oxygen gas at STP = 1/22.414 mole

20 dm³ of oxygen gas at STP = 1/22.414 × 20

= 0.89 mole

2. What volume does 0.6 mole of H₂ gas occupy at STP?

1 mole of H₂ gas occupy a volume = 22.414 dm³ of H₂ at S.T.P.

0.6 mole of H₂ occupy a volume = 0.6 × 22.414 dm³

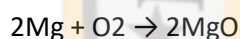
= 13.45 dm³

Q13. How much MgO formed when 4 moles of magnesium react with excess of O₂. Also calculate moles of O₂ required?

Ans. Solution

No. of moles of Mg = 4 moles

According to balanced equation



Similarly

2 moles of Mg produce = 2 mole of MgO

1 moles of Mg require = 1/2 mole of O₂

4 moles of Mg produce = 2/2 × 4 = 4 moles of MgO

4 moles of Mg require = 1/2 × 4 = 2 moles of O₂

Q14. Potassium chlorate is used in making matches and dyes, on decomposition, it produces KCl and oxygen gas. If 25 g KClO₃ is decomposed, calculate the quantities of the following:

(a) How many moles of O₂ are produced. (b) How much mass in gram of O₂ is produced.

Ans. Solution:

(a) How many moles of O₂ are produced

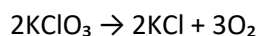
Given mass of KClO₃ = 25 g

Formula mass of KClO₃ = 39 × 1 + 35.5 × 1 + 16 × 3

= 122.5 g mole⁻¹

No. of moles of KClO₃ = 25/122.5 = 0.204 moles

According to balanced equation



2 moles of $\text{KClO}_3 = 3$ moles of O_2

0.204 moles of $\text{KClO}_3 = 3/2 \times 0.204 = 0.306$ moles of O_2

(b) How much mass in gram of O_2 is produced

No. of moles of O_2 produced = 0.306 moles

Molar mass of $\text{O}_2 = 16 \times 2 = 32 \text{ g mole}^{-1}$

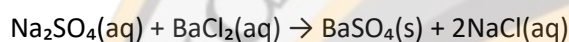
\therefore Mass of O_2 produced = no. of moles of O_2 produced \times molar mass of O_2

= $0.306 \times 32 = 9.792 \text{ g}$

Q15. What Volume of 0.5M sodium sulphate will react with 275 cm^3 of 0.25 M barium chloride solution to completely precipitate Ba^{2+} in solution.

Ans. Solution

The chemical reaction is:



$n_1=1$ $n_2=1$

Thus

Sodium sulphate (Na_2SO_4)

Barium chloride (BaCl_2)

$V_1 = ?$

$V_2 = 275 \text{ cm}^3$

$M_1 = 0.5 \text{ M}$

$M_2 = 0.25 \text{ M}$

$n_1 = 1$

$n_2 = 1$

$$\text{So } \frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2}$$

OR

$$V_1 = \frac{M_2 V_2}{n_2} \times \frac{n_1}{M_1}$$

$$V_1 = \frac{0.25 \times 275}{1} \times \frac{1}{0.5} = 137.5 \text{ cm}^3.$$

Q16. Potassium chlorate (KClO_3) is a white solid. It is used in making matches and dyes. Calculate the molarity of solution that contains.

Ans. (a) 1.5 moles of this compound dissolved in 250 cm^3 of solution

Solution:

No. of moles of $\text{KClO}_3 = 1.5$ moles

Volume of solution = $250 \text{ cm}^3 = 250/1000 \text{ dm}^3 = 0.25 \text{ dm}^3$

Molarity is given by

Molarity = no of moles of solute/volume of solution(dm^3)

Molarity = $1.5 \text{ mol}/0.25 \text{ dm}^3 = 6 \text{ M}$.

(b) 75g of this compound dissolved to produce 1.25 dm³ of solution

Solution:

Given mass of KClO₃ = 75 g

Molar mass of Urea KClO₃ = 39 + 35.5 + 16 × 3 = 122.5 g mol⁻¹

No. of moles of KClO₃ = 75/ 122.5 = 0.612

Volume of solution = 1.25 dm³

Molarity is given by:

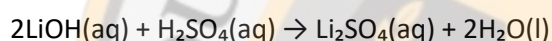
Molarity= no of moles of solute/volume of solution(dm³)

Molarity= 0.612mol/1.25 dm³= 0.49 M.

Q17. What volume of 0.25M LiOH will completely react with 0.500dm³ if 0.25M H₂SO₄

Ans. Solution:

The chemical reaction is:



$n_1=2$ $n_2=1$

Thus

LiOH

H₂SO₄

$V_1 = ?$

$V_2 = 0.5 \text{ dm}^3$

$M_1 = 0.25 \text{ M}$

$M_2 = 0.25 \text{ M}$

$n_1 = 2$

$n_2 = 1$

$$\text{So } \frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2}$$

OR

$$V_1 = \frac{M_2 V_2}{n_2} \times \frac{n_1}{M_1}$$

$$V_1 = \frac{0.25 \times 0.5}{1} \times \frac{2}{0.25} = 1 \text{ dm}^3.$$

Q18. How much volume is occupied by marsh gas (CH₄) at STP containing 4.8×10^{24} molecules.

Ans. Solution:

Since 22.414 dm³ of CH₄ at STP = 1 mole = 6.02×10^{23} molecules.

So

6.02×10^{23} molecules of CH₄ = 22.414 dm³ at STP

4.8×10^{24} molecules of CH₄ =

$$\frac{4.8 \times 10^{24}}{6.02 \times 10^{23}} \times 22.414 = 178.71 \text{ dm}^3 \text{ at STP}$$

Q19. Calculate the molecular weight of a gas with the density of 1.25g/dm³.

Ans. Solution:

Density of gas is 1.25 g/dm³. It means 1 dm³ of gas contains 1.25 g

So

1 dm³ of gas at STP = 1.25 g

22.414 dm³ of gas at STP = 1.25 × 22.414

= 28.02 g

Since

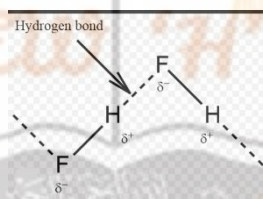
22.414 dm³ of a gas at STP = 1 mole of gas = Molar mass of gas

Therefore, gram molecular mass of gas is 28.02 amu.

Chapter-5: State and Phases of Matter

Q1. What is hydrogen bonding, give particular examples? What are application of H-bonding?

Ans. The electrostatic force of attraction between lone pair of an electron-negative.



Examples: Hydrogen bonding between HF molecules.

Hydrogen bonding has many application.

- Thermodynamic properties of covalent hydrides.
- Solubility of hydrogen bonded molecules.
- Cleansing action of soap & detergents.

Q2. What are the different types of physical properties of liquid? Explain what is (a)Viscosity (b) Surface tension.

Ans. Following are the physical properties of liquid.

Additive Properties: These properties depend upon the number & kind of atoms present in molecule. e.g molecular mass

Constitutive Properties: These properties depend upon the arrangement of atoms in the molecules & not their number. e.g. optical activity

Colligative Properties: These properties depend upon the number of ions & molecules present but do not depend upon the structure of molecules.

e.g. Osmotic pressure etc.

Viscosity: The resistance to flow of a liquid is called "viscosity." The higher the viscosity, the slower the liquid flows.

SI units of viscosity is Pascal Second. Non-SI unit of viscosity is poise.

Surface tension: A force in dynes acting at right angles to a unit length of liquid surface is called "surface tension".

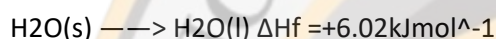
The SI unit of surface tension is joule per square metres. or Newtons per meter 'N/m'

Q3. Define molar heat of fusion & molar heat of vaporization? How will you differentiate crystals from pure liquids?

Ans. Molar heat of Fusion (ΔH_f):

"The amount of heat required to convert one mole of a solid into its liquid state at its melting point."

e.g Molar heat of fusion for ice is +6.02kJ/mole



Molar heat of Vaporization (ΔH_v):

"The amount of heat required to convert one mole of a liquid into its vapours at its boiling point."

e.g. $\text{H}_2\text{O}(l) \longrightarrow \text{H}_2\text{O}(g) \quad \Delta H_v = +40.7 \text{ kJ mol}^{-1}$

Properties of Liquid Crystals/ turbid Phase:

- (1) These have fluidity like liquids.
- (2) These have viscosity, surface tension etc like liquids.
- (3) They have some degree of order like solids.
- (4) They have optical properties like solids.

Q4. Why distillation or vacuum reduced pressure is often used in the purification of chemied? Define dynamic, equilibrium between two physical state?

Ans. "The vacuum distillation liquid boils at low temperature which requires less amounts of heat. In this process boiling point of liquid is decreased."

Where a liquid is placed in a closed container, then molecules leave the surface of liquid & change into vapour state. This is called "evaporation" process. Some of these vapours, on colliding with other vapours, lose a part of their energy & are returned to the liquid state. This is condensation process.

Q5. What are the factors effecting evaporation? Different liquids have different rates of evaporation. Explain with reference to ether & alcohol?

Ans. Factor affecting Evaporation :

Surface Area: "Larger the surface area, higher is the rate of evaporation & vice versa.

Intermolecular force "stronger the intermolecular forces, lower will be the rate of evaporation & vice versa."

Different liquids evaporates a different rates. It is because different strength of inter-molecular forces among their molecules.

Generally, "Stronger the inter-molecular forces, lower will be the rate of evaporation & viceversa."

Alcohol molecules have stronger intermolecular forces between its molecules. The ether molecules have weaker dipole-dipole forces between the molecules. Thus, alcohol has less rate of evaporation than ether.

Q6. Define & explain the term viscosity of a liquid? How does the resistance to the layers causes viscosity? & what are the factors effecting the viscosity of a liquid?

Ans. "The resistance to flow of a liquid is called viscosity." The higher the velocity, the slower the liquid flows.

Unit: SI unit of viscosity is pascal second. Non-SI unit of viscosity is 1poise-

1 poise= $0.1 \text{ kg m}^{-1}\text{s}^{-1}$ or $\text{gcm}^{-1}\text{s}^{-1}$

1 pa.s = $1\text{kg m}^{-1}\text{s}^{-1} = 10 \text{ poise}$.

Factor affecting Viscosity:

Temperature: "Increases in temperature decreases the viscosity of liquid."

"Molecules move faster when the temperature rises. This is because, an increase in the temperature decreases intermolecular forces."

Q7. Define & explain the phenomenon of the surface tension? What are the factors affecting surface tension?

Ans. "The amount of energy required to expand the surface of a liquid per unit area is called surface tension."

Unit: SI unit of surface tension is joule per square meter- Jm^{-2} or Newton per meter, Nm^{-1} .

Explanation: Surface tension is the property of the surface of liquid to behave as if a membrane were stretched over it.

All molecules at the below the surface of a liquid are surrounded by other molecules in all directions. Thus, the force exerted by such molecules is balanced in all directions.

Factor affecting it

(1) Intermolecular Forces: stronger the intermolecular forces, higher the surface tension & vice versa.

(2) Temperature: Increase in the temperature decreases the surface tension & vice versa.

Q8. Evaporation takes place at all temperature. Justify it.

Ans. The heat of sublimation of a substance is greater than its heat of vaporization. Give reason.

Molecules of liquid have K.E at all temperature. Thus they are continuously moving. They collide with one another. As a results, some molecules may get K.E higher than the average value. Such

molecules escape from the liquid surface as vapours. Hence evaporation takes place at all temperature.

The intermolecular attractive forces in solids are stronger than in liquid. Thus, more energy is required to vaporize a solid than to vaporize a liquid. Hence, heat of sublimation is larger than heat of vaporization.

Q9. The change in the surface area of a liquid does not cause the change in the equilibrium vapour pressure of liquid? how? explain why the steam at 100°C burns your skin more severely than liquid water at 100°C?

Ans. "The evaporation is faster from large surface area | however, remaining molecules 'also have large area to retain. This condensation is also fast."

When a liquid is heated, its temperature goes on increasing until its boiling point is reached. At boiling point temperature becomes constant. It is because the heat supplied at boiling point is used to convert the liquid into vapour by breaking intermolecular forces.

Q10. Indicate the effect of pressure & temperature on solids? If kinetically how will you interpret the effect of temperature on gaseous molecules?

Ans. These are little affected by changes in the temperature & pressure.

In solids, molecules are tightly packed, therefore, pressure has no effect on them. An increase in temperature increases the volume of a gas.

According to kinetic molecular theory of gases, The average K.E. of gas molecules is directly proportional to the absolute temperature i.e. $K.E \propto T$.

Q11. Name four elements & four compounds that exist as gases at room temperature? & Differentiate b/w amorphous solids & crystalline solids?

Ans. Elements as Gases: He, Ne, Ar, H₂, N₂, O₂ etc

Compounds as Gases: CH₄, CO₂, CO, NH₃ etc

Amorphous	Crystalline
The solids which have no definite geometric shape.	The solid which has a definite regular & three-dimensional geometric solid/shape. e.g. NaCl has cubic shape.
They do not have the regular arrangement of atoms, ions or molecules. Therefore, they may be called as super-cooled liquids.	They have complete regular arrangement of atoms, ions & molecules.
They do not possess sharp melting point. They are softened on heating.	They possess sharp melting point.

Chapter 6: Energetics

Q1. Define the terms system, surrounding boundary, state function, heat capacity, internal energy, enthalpy in substance?

Ans. Systems: Any real or imaginary part of the universe that is under study is called "system". A chemical system is usually a substance undergoing a chemical change.

Surroundings: Everything that is not part of system is called surrounding.

Boundary: The real or imaginary surface separating the system from surrounding is called "Boundary".

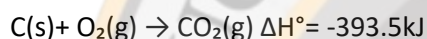
State Function: It is a microscopic property of system which has definite values for initial & final states & it is independent of the path through which change take place. Examples: volume, Temperature, Enthalpy.

Internal Energy: It is the sum of all the energies of all the particle within a system. Denoted by 'E'. Given as $E=K.E+P.E$

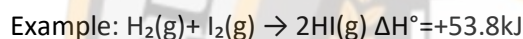
Q2. Classify reaction as exothermic or endothermic? & Define bond dissociation energy?

Ans. Exothermic Reaction: A chemical reactions that proceeds with the evolution of heat.

Example: Burning of fuels is a highly exothermic reaction.



Endothermic Reaction: "A chemical reactions that proceeds with absorption of heat."



Dissociation Energy: The energy required when one mole of a particular bond breaks apart to form neutral atoms are called the bond dissociation energy.

Bond Energy: The amount of energy released when one mole of neutral atoms form bond is called bond energy."

Q3. When ethanol burns in oxygen, carbon dioxide & water are formed

(a) Write the equations which describes this reaction?

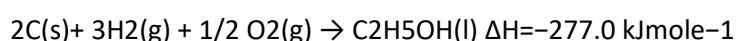
(b) Using the following data, calculate the enthalpy of combustion for ethanol C_2H_5OH .

ΔH_f for ethanol (l) = $-277.0 \text{ kJmole}^{-1}$

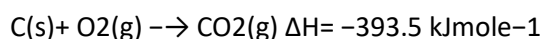
$\Delta H_f CO_2(g) = -393.5 \text{ kJmole}^{-1}$, $\Delta H_f \text{ water}(l) = -285.8 \text{ kJmole}^{-1}$

Ans. Solution:

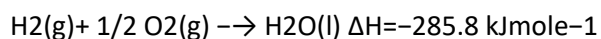
Given Data: Heat of formation of liquid ethanol $C_2H_5OH(l)$



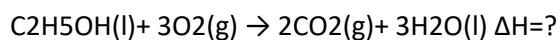
Heat formation of gaseous $CO_2(g)$



Heat of formation of liquid H_2O

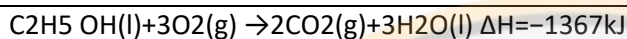
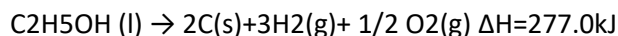
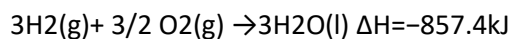
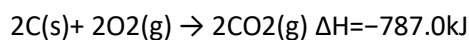


Required: Heat of combustion of liquid $C_2H_5OH(l)$



Multiply Eq. 2 by 2, Multiply Eq. 3 by 3 Reverse Eq. (1)

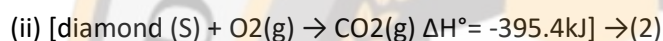
Then add the resulting equations, we get



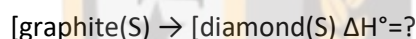
Q4. Calculate ΔH for the conversion of graphite to diamond. Heats of combustion of graphite & diamond measure at 25°C & 1 atm pressure are -393.5 kJ resp -395.4 kJ respectively & Develop a hypothetical reaction & calculate the Gibbs free energy Discuss conditions under which reaction would be spontaneous or non-spontaneous.

Ans. Solutions:

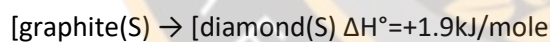
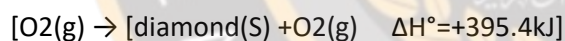
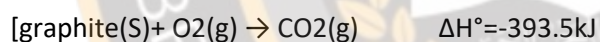
Given Data:



Required:

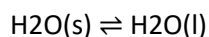


Reverse the Eq. (2) & add it to Eq. 1, we get-



Gibbs free energy is useful for understanding phase transition, such as a form liquid to solid or gas & vice versa.

Example: Consider, the melting of ice



At the melting point, $\Delta G = 0$, So, the solid & liquid phases are in equilibrium.

Q5. Write the names of factors that affect enthalpy of hydration? Diff enthalpy of hydration from lattice energy?

Ans. (1) Ionic Charge: Increase in ionic charge increases the enthalpy of hydration & vice versa.

(2) Ionic Radius: Increase in ionic radius decreases the enthalpy of hydration & vice versa.

Hydration	Lattice
Is the enthalpy change when one mole of solid dissolves in water?	It is enthalpy change for the conversion of crystalline ionic solid

	into gaseous ions vice versa.
This process involves solid & liquid (aq.) phases.	This process involves solid & gas phase.
It is related to the solubility of a substance.	It is related to the stability of a substance.

Q6. Why it is essential to mention the physical states of reactants & products in a thermochemical equation? Enthalpy change is state function but heat is not?

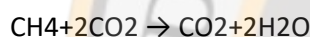
Ans. Substances have different energies in different physical states. Therefore, heat changes will be different if or when same the substance is involved is different physical states.

Heat is always the energy in transfer, i.e the energy which 'crosses' the system - boundaries. Hence, It depends upon path & it is path function. Enthalpy is the total heat content in a system.

Q7. Can standard enthalpy of reaction(ΔH°) be negative? & Heat of combustion is always negative?

Ans. In exothermic reaction the total enthalpy of the product is less than those of the reactant. Thus, for exothermic reaction ΔH , is negative. e.g Combustion of methane in oxygen is an exothermic process.

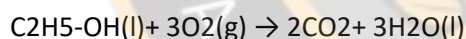
e.g.



$$\Delta H^\circ = -890.4 \text{ kJ mol}^{-1}$$

Combustion reaction are always exothermic. In this reaction energy of reactant is always greater than that of product. Hence, the energy will be released & the enthalpy change will always be negative.

e.g.



$$\Delta H^\circ = -1367 \text{ kJ mol}^{-1}$$

Q8. What is entropy?

Ans. The degree of disorder and randomness of a system is called its entropy. It is denoted with the symbol "S". As a system moves from a more ordered state to a less ordered state, the entropy of the system tends to increase. Melting ice gives us a great example of entropy.

The entropy change of the system is denoted by ΔS .

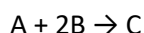
- If the entropy of the system increases, this is represented as a positive change $\Delta S > 0$
- When the entropy of a system decreases, it is represented as a negative change $\Delta S < 0$.
- In general, when the temperature of a system increases, its entropy also increases, its $\Delta S > 0$.
- As the temperature of the system decreases, its $\Delta S < 0$.

Calculation of The Entropy Change

The entropy change for a chemical reaction can be calculated by the following equation.

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

Suppose we want to calculate ΔS° for the following reaction.



$$S^\circ(A) = 40\text{kJ/mol.}$$

$$S^\circ(B) = 50\text{kJ/mol.}$$

$$S^\circ(C) = 100\text{kJ/mol.}$$

Solution

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$= [1(100)] - [1(40) + 2(50)]$$

$$= 100 - [40 + 100]$$

$$= -40\text{kJ/mol}$$

Q9. Define a liquid crystal? What are the uses of liquid crystals in daily life?

Ans. The liquid crystal is intermediate between solid crystal and clear liquid its properties are also between crystals properties and those of isotropic liquids.

Uses of liquid crystals

1. Nematic and cholesteric Liquid crystals have electromagnetic optic characteristic and photoelectric properties. Thus they been used in different laboratories, universities and industries began to focus on their applications.
2. Liquid crystals are used in the display of electrical devices like digital watches, laptops computers, calculators etc.
3. Cholesteric liquid crystals applied to skin are used to locate veins, arteries, infections, tumours etc. These parts are warmer than other tissues. So, liquid crystals change their colour.
4. Nematic liquid crystals are used in magnetic resonance. Molecules dissolved in nematic crystal liquid solvents show resolved NMR spectrum. Intermolecular Dipole-Dipole fine structures are obtained.
5. Liquid crystals are used in chromatographic separation to study molecular arrangements, kinetics. These are also used as anisotropic fluid for visible, UV and IR spectroscopy.
6. Liquid crystals are used in Oscillography system, television displays, liquid crystal screens.

Q10. Classify the following processes as exothermic or endothermic.

Ans. (a) Freezing of water : During freezing process, the temperature is decreased. So, heat is released from the system. It is an exothermic process.

(b) Combustion of methane: During combustion process, heat is evolved. So it is an exothermic process.

(c) Sublimation of dry ice: Sublimation of ice involves the conversion of ice molecules into vapours. For this heat is absorbed. So, it is an endothermic process.

(d) $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$: In this water vapours ($\text{H}_2\text{O}(\text{g})$) are converted into liquid water ($\text{H}_2\text{O}(\text{l})$). So, energy is released. It is an exothermic process.

(e) Decomposition of limestone: The limestone is decomposed on heating. So heat is absorbed. It is an endothermic process.

Q11. Define the following.

Standard Enthalpy of Formation: The enthalpy change that accompanies the formation of one mole of a compound from its elements with all substances in their standard states.

Standard Enthalpy of Combustion ($\Delta H^\circ\text{c}$): The enthalpy change when one mole of a substance is completely burnt in excess of oxygen under standard conditions is called standard enthalpy of combustion.

Standard Enthalpy of Atomization ($\Delta H^\circ\text{at}$): The enthalpy change when one mole of gaseous atoms are formed from its element under standard conditions is called standard enthalpy of atomization

Standard Enthalpy of Solution: It is the enthalpy change when one mole of a substance is dissolved in so much of solvent that further dilution results in no detectable heat change, under standard conditions

Lattice Energy: The energy released when one mole of an ionic compound (solid) is formed from its constituent gaseous ions is called lattice energy.

Chapter-7: Chemical Kinetics

Q1. How do you differentiate chemical kinetics from chemical equilibrium? & Explain effect of concentrations, temperature & surface area on reaction rates?

Ans. Chemical Kinetics: It is study of rates of chemical reaction & the factors that affect the rates of chemical reactions.

It gives information about mechanism of reaction.

Chemical equilibrium: It is state of a reversible reaction at which the rate of forward reaction becomes equal to the rate of backward reaction.

It is only concerned with reversible reaction.

Concentration of Reactant: The reaction b/w concentration & the rate of reaction is given by the law of Mass Action.

It states that rate of chemical reaction is proportional to the product of molar concentrations of reactants"

Temperature: Reaction rates generally increase with the increase in temperature.

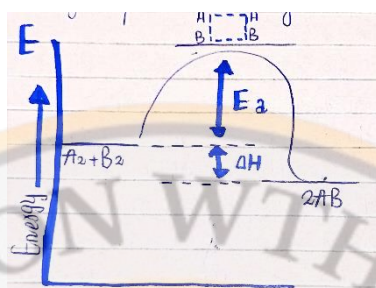
Surface Area: The rate of reaction increases with increasing surface area of reactants.

Q2. Evaluate that increase in collision energy by increasing the temperature can improve the collision frequency? & Draw energy diagram that represent the activation energy & show effect of catalyst.

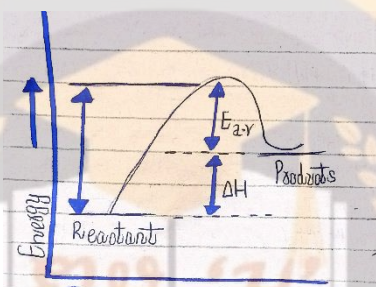
"The reaction rate generally increases with the increasing temperature."

According to the collision theory, the rate of a reaction is proportional to the number of collision among the reactant molecules.

An increase in temperature increase the average kinetic energy of molecules. This increases average speed of reading molecules.



Reaction path -> (a)



Reaction path -> (b)

(a) Exothermic reaction

(b) Endothermic rxn

Q3. What is the effect of a catalyst on the following?

(i) The rate of reaction

(ii) The energy of activation.

The reaction of an alkyl halide, R-X with water is as follows:



If the reaction were of single step then what would be the rate law?

Ans. (i) The rate of Reaction: A catalyst provides a new mechanism for the reaction with low energy of activation.

(ii) The energy of activation:

"A catalyst provides a new mechanism for the reaction with low energy of activation.

Since, the reaction is a single step process, therefore, this single step is the rate determining step. In this step- One molecule of R-X & one molecule H₂O are reacting together.

Q4. A very small amount of catalyst may prove sufficient to carry out reaction? & the reaction rate decreases every moment?

Ans. A catalyst is not consumed during a reaction. It is recycled again & again. Thus, only small amount catalyst is sufficient to carry out reaction.

Example:

- 1 mg of fine platinum powder can convert 2.5 dm³ of H₂ & 1.25 dm³ of O₂ to water.

According to the law of mass actions rate of a chemical reaction is directly proportional to concentration of reactant. When the reaction starts, the concentration of reactants is high, therefore rate of reaction is fast. As the concentration of reactants is decreased with time, the rate of a reaction is also decreased.

Chapter-8: Chemical Equilibrium

Q1. Define & explain the following terms?

- (i) Macroscopic events
- (ii) Dynamic equilibrium
- (iii) Le chatlier's principal
- (iv) Heterogenous equilibrium
- (v) Homogenous equilibrium

Dynamic Equilibrium: "The stage of a reversible reaction at which rate of forward reaction becomes equal to rate of the reverse reaction."

Homogeneous Equilibrium: "It is the type of equilibrium in which all the reactants & products are presents in the same phase."

Heterogeneous Equilibrium: "It is the type of equilibrium in which all the reactants & products are present in more than one phase."

Le-Chatlier's Principle:-"It states: If a change is imposed on a system at equilibrium, the position of the equilibrium will shift in a direction which tends to reduce that change."

Microscopic events: Microscopic events refer to phenomena that cannot be observed with the naked eye.

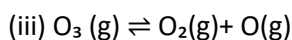
Q2. Write K_c & K_p expressions for following reaction?

Ans. (i) $\text{SO}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightleftharpoons \text{SO}_3(\text{g})$

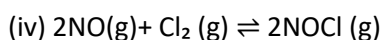
$$K_c = \frac{[\text{HOCL}]^2}{[\text{H}_2\text{O}][\text{Cl}_2\text{O}]} \quad \& \quad K_p = \frac{[\text{PSO}_3]}{P_{\text{SO}_2} \times P_{\text{O}_2}^{1/2}}$$

(ii) $\text{H}_2\text{O}(\text{g}) + \text{Cl}_2\text{O}(\text{g}) \rightleftharpoons 2\text{HOCl}(\text{g})$

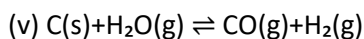
$$K_c = \frac{[\text{SO}_3]}{[\text{SO}_2][\text{O}_2]^{1/2}} \quad K_p = \frac{p^2 \text{HOCL}}{P_{\text{H}_2\text{O}} \times P_{\text{Cl}_2\text{O}}}$$



$$K_c = \frac{[O_2][O]}{[O_3]} \quad K_p = \frac{P_{O_2} \times P_O}{P_{O_3}}$$



$$K_c = \frac{[NOCl]^2}{[NO]^2[Cl_2]} \quad K_p = \frac{p^2 NOCl}{P^2 NO \times P_{Cl_2}}$$



Since Solids are not included in equilibrium constant expressions, therefore

$$K_c = \frac{[CO][H_2]}{[H_2O]} \quad K_p = \frac{P_{CO} \times P_{H_2}}{P_{H_2O}}$$

Q3. At a particular temperature $K_p = 0.133$ atm.

Which of the following conditions corresponds to equilibrium position for the reaction $A(g) \rightarrow 2B(g)$

(a) $P_B = 0.175$ atm, $P_A = 0.102$ atm

(b) $P_B = 0.064$ atm, $P_A = 0.0308$ atm

(c) $P_B = 0.144$ atm, $P_A = 0.156$ atm

Ans. Solution:

$K_p = 0.133$ atm

The reaction quotient is determined as follows

$$Q = \frac{P^2 B}{P_A}$$

The value of Q is compared with K_c . When these two values are equal then reaction will be at the equilibrium. For the given reaction, the reaction quotient is-

$$Q = \frac{P^2 B}{P_A}$$

(a) $P_B = 0.175$ atm, $P_A = 0.102$ atm

$$Q = \frac{P^2 B}{P_A} = \frac{(0.175)^2}{0.102} = 0.30 \text{ atm.}$$

Since, $Q > K_p$ therefore, reaction is not at equilibrium.

It will proceed in backward direction.

(b) $P_B = 0.064$ atm, $P_A = 0.0308$ atm

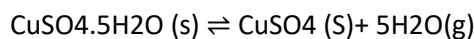
$$Q = \frac{P^2 B}{P_A} = \frac{(0.064)^2}{0.0308} = 0.133 \text{ atm.}$$

Since $Q = K_p$, therefore, reaction is at equilibrium

Q4. Write the expression for Kc & Kp for the following processes.

(i) Blue vitriol is deep blue solid copper (II) sulphate pentahydrate is heated to drive off water vapour to form white solid copper (II) sulphate.

Ans. The chemical reaction is



Since, Solids are not included in equilibrium constant expression, therefore

$$K_c = [\text{H}_2\text{O}]^5 \quad \& \quad K_p = p^5 \text{H}_2\text{O}$$

(ii) The decomposition of solid phosphorous pentachloride to gaseous phosphorous trichloride & chlorine gas

The chemical reaction is



Since, Solids are not included in equilibrium constant expression, therefore

$$K_c = [\text{PCl}_3][\text{Cl}_2] \quad \& \quad K_p = P_{\text{PCl}_3} \times P_{\text{Cl}_2}$$

Q5. What is difference between an equilibrium with a Kc value larger than one compared with an equilibrium that has Kc smaller than one?

Ans. Solution:

For a general reaction

Reactants \rightleftharpoons Products

The Kc is given as

$$K_c = \frac{[\text{products}]}{[\text{reactants}]}$$

If $K_c > 1$, The equilibrium mixture contains more product than reactant. Thus Forward reaction occurs to greater extent.

If $K_c < 1$, the equilibrium mixture contains more reactant than products. Thus backward reaction occurs to greater extent.

If $K_c = 1$, the equilibrium mixture contains almost same amount of product & reactants.

Q6. For some reaction the four types of chemical equilibrium constant i.e, Kc, Kp, Kx & Kn become equal: Which of following reaction involve homogeneous equilibriums or which involve heterogeneous equilibrium.

Ans. The relationship b/w different equilibrium constants is

$$K_p = K_c (RT)^{\Delta n} = K_x (P)^{\Delta n} = K_n (P/N)^{\Delta n}$$

If the number of moles of reactants & products are equal for a reaction, or $\Delta n = 0$

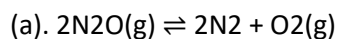
Then,

$$(RT)^0 = 1, (P)^0 = 1, [P/N] = 1$$

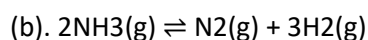
So,

$$K_p = K_c = K_x = K_n$$

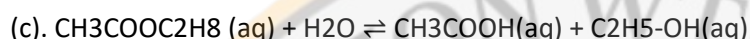
It means that when the numbers of moles of reactants & products in a reaction are equal, then all equilibrium constant have same values.



Since, all component are gases i.e is same phase. So, it is homogenous equilibrium.



Since, all component are gases. i.e in same phase. So, it is homogenous equilibrium.



Since, all component are in solution i.e in the same aqueous phase. So, it is homo-geneous equilibrium.

Q7. The solubility of glucose in water is increased by increasing the temperature. Different b/w reversible & irreversible reactions?

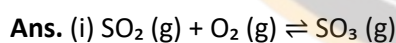
Ans. The solubility of glucose in water is an endothermic process



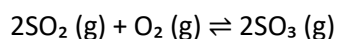
So, according to Le-Chatliers Principle, if temprature is increased, the system will move in forward direction.

Reversible	Irreversible
The reaction which can proceed both in the forward as well as in backward direction.	The reaction which take place only in one directions, are called "irreversible reaction."
In these reactions, dynamic equilibrium is established.	In these reactions, one or more reactants are almost consumed completely.
Examples: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$	Example: $2Na + 2H_2O \rightarrow 2NaOH + H_2$

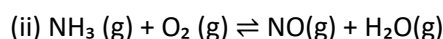
Q8. The following equations represent various industrial reactions at equilibrium. Write Kc expressions in mole fractions, of these reactions. Do not forget to balance the equations:



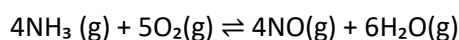
The balanced equation is



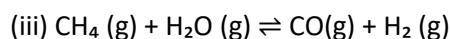
$$K_c = \frac{[SO_3]^2}{[SO_2]^2 [O_2]}$$



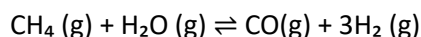
The balanced equation is



$$K_c = \frac{[NO]^4 [H_2O]^6}{[NH_3]^4 [O_2]^5}$$



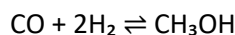
The balanced equation is



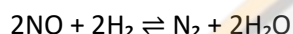
$$K_c = \frac{[\text{CO}][\text{H}_2]^3}{[\text{CH}_4][\text{H}_2\text{O}]}$$

(b) Give the balanced equations that correspond to following equilibrium expressions.

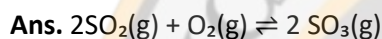
(i) $K_c = \frac{[\text{CH}_3\text{OH}]}{[\text{CO}][\text{H}_2]^2}$



(ii) $K_c = \frac{[\text{N}_2][\text{H}_2\text{O}]^2}{[\text{NO}]^2[\text{H}_2]^2}$



Q9. The contact process prepares purest sulphuric acid commercially. Following reaction takes place in the contact chamber in the presence of V_2O_5 .



Calculate K_p if the following concentrations are found at equilibrium at 25°C .

$$[\text{SO}_2] = 0.59\text{M}, [\text{O}_2] = 0.05\text{M} \text{ \& \ } [\text{SO}_3] = 0.259\text{M}$$

Calculation of K_c

K_c is given as

$$K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]}$$

$$K_c = \frac{(0.259)^2}{(0.59)^2(0.05)}$$

$$K_c = 3.85$$

Calculation of K_p

$$\Delta n = 2 - (2 + 1) = -1$$

$$R = 0.0821 \text{ atm dm}^3 \text{ K}^{-1} \text{ mole}^{-1}$$

$$T = 25^\circ\text{C} + 273 = 298\text{K}$$

$$K_c = 3.85$$

Thus

$$K_p = K_c (RT)^{\Delta n}$$

$$K_p = 3.85 \times (0.0821 \times 298)^{-1}$$

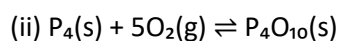
$$K_p = 3.85 / (0.0821 \times 298)$$

$$K_p = 0.1574$$

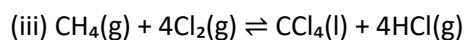
Q10. Write K_c and K_p expressions for each of the following reactions.



$$K_c = [\text{CO}_2]/[\text{CO}] \quad \& \quad K_p = P_{\text{CO}_2}/P_{\text{CO}}$$



$$K_c = [\text{O}_2]^{-5} \quad \text{and} \quad K_p = P_{\text{O}_2}^{-5}$$



$$K_c = [\text{HCl}]^4/[\text{CH}_4] [\text{Cl}_2]^4 \quad \text{and} \quad K_p = P_{\text{HCl}}^4/P_{\text{CH}_4} \times P_{\text{Cl}_2}^4$$

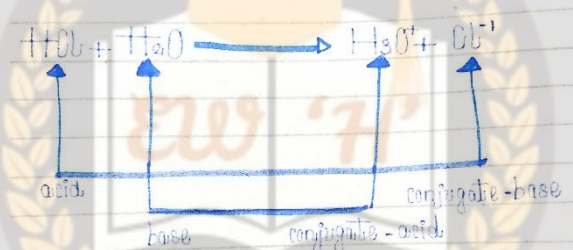
Chapter 9: Acid-Base Chemistry

Q1. What are conjugate acid-base pairs? Explain with example. & Define Lewis acid & Lewis base. Give one example in each case?

Ans. The conjugate acid-base pairs are species on opposite sides of an equation that differ by a proton. The weaker acids have strong base pairs & stronger acids has weaker conjugate bases.

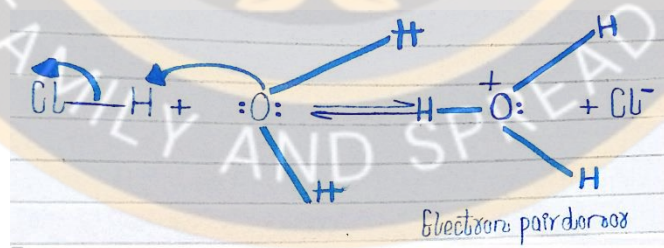
Examples:

(i) Ionisation of HCl in water



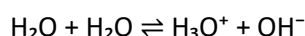
"A Lewis acid is substance that can accept a pair of electron to form a coordinate covalent bond."

"A Lewis base is a substance that can donate a pair of electrons to form a coordinate covalent bond."



Q2. Write briefly about the ionisation of water ? & Define PH what are the values of PH for acidic, basic & neutral solutions?

Ans. Water is a unique compound due to its ability to accept or donate protons under blue different environments. Water itself undergoes ionization to a small extent as shown in following equation.



PH: The negative logarithm of H^+ ions concentration (in mol dm^{-3}) is called PH.

Acid solutions: $[\text{H}^+] > 1.0 \times 10^{-7}\text{M}$, $\text{PH} < 7.00$

Basic solution: $[\text{H}^+] < 1.0 \times 10^{-7}\text{M}$, $\text{PH} > 7.00$

Neutral solution: $[H^+] = 1.0 \times 10^{-7} M$, $PH = 7.00$

POH: The negative logarithm of OH ions concentration (in mol dm^{-3}) is called POH.

$pOH = -\log[OH^-]$

Q3. What are Ka & pKa? & What are Kb & pKb?

Ans. Ka is called the acid dissociation constant.

It is a measure of extent to which an acid is ionized, or dissociated at equilibrium state.

PKa: pKa is defined as the negative logarithm of Ka.

$$pKa = -\log Ka$$

Kb is also known as dissociation constant of a base.

PKb: PKb is defined as the negative logarithm of Kb.

$$PKb = -\log Kb$$

Q4. What is the relationship b/w Ka & Kb? & Give two examples of buffer solutions?

Ans. A strong acid produces a relatively weak conjugate base.

A strong base produces a relatively weak conjugate acid.

Example: CH_3COOH is a weak acid.

HNO_3 is a strong acid.

Acidic buffer: A buffer solution may contain a weak acid & the salt of that acid with strong base.

e.g. $H_2CO_3 + NaHCO_3$

Basic buffer: A buffer solution may contain a weak base & its salt with a strong acid.

e.g. $NH_4OH + NH_4Cl$

Such solution will give basic buffers with PH more than 7.

Calculations of PH of Buffer solutions:

Since salt is a strong electrolyte. It is ionized almost completely. Thus, the concentration of the conjugate base in sxn mixture is assumed equal to that of salt.

Q5. Give two examples of a buffer solution? How does temperature impact solubility?

Ans. Thus, pH of a buffer can be calculated by using Hendersons -Hasselbalch equation.

$PH = Pka + \log \frac{[Salt]}{[Acid]}$

- Increase in temperature may increase the solubility of substance. In such cases, ΔH_{sol} is usually positive. (Endothermic)
- Increase in temperature may decrease the solubility of substance. In such cases, ΔH_{sol} is usually negative (Exothermic).
- However, there are certain substances for which ΔH_{sol} is small

Q6. (a) Calculate the PH of formic acid- sodium formate buffer solutions containing 1.0 mole of each component? & (b) Calculate the H⁺ ions concentration of an aqueous solution having PH 10.6

Ans. Formic acid is HCOOH & sodium formate is HCOONa, Assuming volume of buffer solution is 1dm³, we have

$$[\text{HCOONa}] = 1.0 \text{ mole dm}^{-3}$$

$$[\text{HCOOH}] = 1.0 \text{ mole dm}^{-3}$$

$$\text{Ka of formic acid} = 1.8 \times 10^{-4}$$

$$\text{So pKa} = -\log \text{Ka}$$

$$= -\log(1.8 \times 10^{-4})$$

$$= 3.744$$

Hence, according to Henderson's eq. pH of buffer is

$$\text{pH} = \text{pKa} + \log [\text{salt}]/[\text{acid}]$$

$$\text{pH} = \text{pKa} + \log [\text{HCOONa}]/[\text{HCOOH}]$$

$$\text{pH} = 3.744 + \log [1.0]/[1.0]$$

$$\text{pH} = 3.744 + 0$$

$$\text{pH} = 3.744$$

$$\text{Solution: pH} = 10.6 \text{ } [H^+] = ?$$

As:

$$\text{pH} = -\log [H^+]$$

$$10.6 = -\log [H^+]$$

or

$$\log [H^+] = -10.6$$

Taking antilog on both sides

$$[H^+] = \text{antilog} [-10.6]$$

$$[H^+] = 2.51 \times 10^{-11} \text{ mol/dm}^3$$

Q7. An aqueous solution contain 1.0×10⁻⁹ moles/dm³ of hydronium ions. Calculate the pOH of this solution. So What is hydrolysis? Write the equation of hydrolysis equilibrium for each of the followings?

Ans. Solution:

Hydronium ion (H₃O⁺) represents the H⁺ ions in aqueous solution, SO

$$[H^+] = 1.0 \times 10^{-9}$$

$$\text{POH} = ?$$

As:

$$PH = -\log [H^+]$$

$$PH = -\log (1.0 \times 10^{-9})$$

$$PH = -\log (10^{-9})$$

$$PH = -(-9) \log(10)$$

$$PH = 9 \times 1 = 9$$

Since:

$$PH + POH = 14$$

$$9 + pOH = 14$$

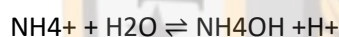
$$pOH = 14 - 9$$

$$pOH = 5$$

i) K^+ : Since, K^+ is derived from relatively a stronger base. Therefore, it will not hydrolyse with water.

NH_4^+ : NH_4^+ is a strong conjugate acid of weak base NH_4OH . According to Bronsted-Lowry Concepts contains react with water producing acidic solutions .

Thus.



Q8. Which of the following can act as Bronsted acid & which as Bronsted base? Point out the conjugate acid base pairs in following reversible reaction?

Ans. (a) H_3O^+ : H_2O donates proton, so it is a Bronsted acid :: $H_3O^+ + OH^- \rightleftharpoons H_2O + H_2O$

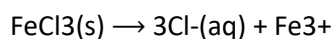
(b) Cl^- : Cl^- accepts protons, so it is a Bronsted base:- $Cl^- + H^+ \rightleftharpoons HCl$

(c) $HNO_3 + H_2O \rightleftharpoons H_3O^+ + NO_3^-$

Since, HNO_3 donates a proton to form the NO_3^- . So, HNO_3 is acid & NO_3^- is its conjugate base.

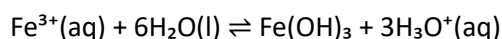
Q9. Why is the solution of ferric chloride acidic in water? $AlCl_3$ is said to be Lewis acid but not Bronsted -lowery acid. why?

Ans. This salt is formed from a strong acid (HCl) & a weak base [$Fe(OH)_3$]. It ionizes as H_2O (excess)



or The anion (Cl^-) coming from strong acid does not react further with water.

The cation (Fe^{3+}) coming from weak base reacts with water & thus undergoes hydrolysis as:



Form salt weak base free hydronium ion

In $AlCl_3$, the octet of Al is not complete. It has six electrons in its valence shell. Thus, it is an electron-deficiency molecule. So it can accept an electron pair. According to the Lewis concept, electron pair acceptors are called acids. Hence $AlCl_3$ is Lewis acid.

Q10. The concentration of [OH⁻] ion in a household ammonia solution is 0.005M. Calculate the concentration of [H⁺] in it.

Ans. Solution:

$$[\text{OH}^-] = 0.005 \text{ M}$$

$$K_w = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

Thus

$$K_w = [\text{H}^+][\text{OH}^-]$$

$$1.0 \times 10^{-14} = [\text{H}^+] \times 0.005$$

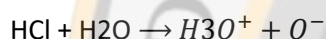
$$[\text{H}^+] = 1.0 \times 10^{-14} / 0.005$$

$$= 2.0 \times 10^{-12} \text{ M}$$

Q11. Calculate the pH of 0.001 M aqueous hydrochloric acid solution.

Ans. Solution:

Hydrochloric acid ionises in water completely therefore,



$$0.001 \text{ M} \quad 0.001 \text{ M}$$

H_3O^+ is in fact the same as $[\text{H}^+]$

$$\text{Therefore } [\text{H}^+] = 0.001 \text{ M}$$

$$\text{pH} = -\log [\text{H}^+]$$

$$\text{pH} = -\log (0.001)$$

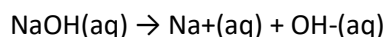
$$= -\log 10^{-3}$$

$$= 3.00$$

Therefore, the pH of 0.001 M aqueous hydrochloric acid is 3.00

Q12. Calculate the pH of 0.062 M NaOH solution

Ans. Solution:



$$0.062\text{M} \quad 0.062\text{M} \quad 0.062\text{M}$$

$$[\text{OH}^-] = 0.062 \text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-]$$

$$= -\log [0.062] = 1.21$$

$$\text{As } \text{pH} + \text{pOH} = 14$$

$$\text{So } \text{pH} = 14 - \text{pOH}$$

$$= 14 - 1.21 = 12.79$$

Q13. What is the pH of a solution containing 1.95g pure H₂SO₄, per dm³ of solution?

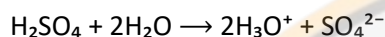
Ans. Solution:

$$\text{Mass of H}_2\text{SO}_4 = 1.95 \text{ g}$$

$$\text{Molecular mass of H}_2\text{SO}_4 = 1 \times 2 + 32 + 16 \times 4 = 98 \text{ g mole}^{-1}$$

$$\text{Number of moles of H}_2\text{SO}_4 = 1.95 / 98 = 0.02 \text{ moles}$$

H₂SO₄ ionises in water completely therefore,



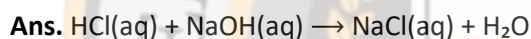
$$0.02 \qquad \qquad \qquad 2 \times 0.02$$

[H₃O⁺] is in fact the same as [H⁺]

$$\text{Therefore } [\text{H}^+] = 2 \times 0.02 = 0.04 \text{ M}$$

$$\text{So pH} = -\log(0.04) = 1.40$$

Q14. In a titration it is found that 25 cm³ of 0.12M NaOH is neutralized with 30cm³ of HCl of unknown concentration. Calculate concentration and strength of HCl solution.



$$n_1 = 1 \quad n_2 = 1$$



$$M_1 V_1/n_1 = M_2 V_2/n_2$$

$$M_1 \times 30/1 = 0.12 \times 25/1$$

$$M_1 = 0.1\text{M}$$

Thus Molarity of HCl solution is 0.1M.

$$\text{Strength of solution} = \text{Molarity} \times \text{Molar mass}$$

$$\text{Strength of HCl solution} = 0.1 \times 36.5$$

$$= 3.65 \text{ g dm}^{-3}$$

Q15. What is the pH of buffer if concentration of CH₃COOH is 0.1 M and CH₃COONa is 1.0 M pK_a for CH₃COOH is 4.76?

Ans. Solution:

$$\text{Concentration of CH}_3\text{COOH} = 0.1 \text{ M}$$

$$\text{Concentration of CH}_3\text{COONa} = 1.0 \text{ M}$$

The formula to determine pH value is

$$\text{pH} = \text{pK}_a + \log [\text{salt}] \text{ (i.e. CH}_3\text{COONa)} / [\text{acid}] \text{ (i.e. CH}_3\text{COOH)}$$

$$\text{pH} = 4.76 + \log (1.0/0.1)$$

$$\text{pH} = 4.76 + \log_{10} (1.0/0.1 = 10/1)$$

$$\text{pH} = 4.76 + 1.00 (\log 10 = 1)$$

$$\text{pH} = 5.76$$

Q16. Calculate the pH of a buffer solution in which 0.11 Molar CH_3COONa and 0.09 Molar CH_3COOH solutions are present. K_a for CH_3COOH is 1.8×10^{-5}

Ans. Solution

Concentration of $\text{CH}_3\text{COOH} = 0.09 \text{ M}$

Concentration of $\text{CH}_3\text{COONa} = 0.11 \text{ M}$

K_a of acetic acid = 1.8×10^{-5}

So, $\text{p}K_a = -\log K_a$

$$= -\log (1.8 \times 10^{-5})$$

$$= 4.745$$

The formula to determine pH value is

$$\text{pH} = \text{p}K_a + \log [\text{salt}] (\text{i.e. } \text{CH}_3\text{COONa}) / [\text{acid}] (\text{i.e. } \text{CH}_3\text{COOH})$$

$$\text{pH} = 4.745 + \log (0.11/0.09)$$

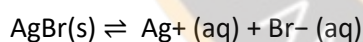
$$\text{pH} = 4.745 + 0.087$$

$$\text{pH} = 4.832$$

Q17. The solubility of Ag Br is $7.1 \times 10^{-7} \text{ M}$ at 25°C . Calculate its K_{sp} .

Ans. Solution

$7.1 \times 10^{-7} \text{ M}$ dissolved Ag Br produces equal moles of Ag^+ and Br^- ions.



$$7.1 \times 10^{-7} \text{ M} \quad 7.1 \times 10^{-7} \text{ M}$$

Thus

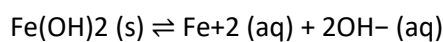
$$K_{sp} = [\text{Ag}^+][\text{Br}^-]$$

$$K_{sp} = (7.1 \times 10^{-7}) (7.1 \times 10^{-7})$$

$$K_{sp} = 5.041 \times 10^{-13}$$

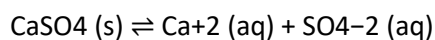
Q18. (i) Write K_{sp} expressions for a) Iron (II) Hydroxide b) Calcium Sulphate

Ans. (a) Iron(II)Hydroxide



$$K_{sp} \text{ is given as } K_{sp} = [\text{Fe}^{2+}] [\text{OH}^-]^2$$

(b) Calcium Sulphate

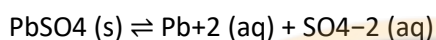


Ksp is given as $K_{sp} = [\text{Ca}^{+2}] [\text{SO}_4^{-2}]$

(ii). Lead (II) Sulphate is used as white pigment. What is the solubility of PbSO_4 ?

$K_{sp} = 1.96 \times 10^{-8}$ at 25°C

Let x is the solubility of PbSO_4 . Let an excess amount 'a' of PbSO_4 is mixed with water then PbSO_4 ionizes in water as



Initial Conc. a 0 0

(in moles)

Conc. at a - x x x

Equilibrium

(in moles)

KSP is given as

$$K_{SP} = [\text{Pb}^{+2}] [\text{SO}_4^{-2}]$$

$$\text{or } K_{SP} = (x) (x) = 1.96 \times 10^{-8}$$

$$x^2 = 1.96 \times 10^{-8}$$

Taking square root on both sides

$$x = \sqrt{1.96 \times 10^{-8}}$$

$$x = 1.40 \times 10^{-4} \text{ M}$$

Chapter 10: Periodic Table

Q1. Differentiate between S and p block elements & find out the position of sulphur in periodic table?

S-block	P-block
They have valence electrons in s-subshell.	They have valence electron in p-subshell.
They include metal.	They include metals, non-metal & metalloids
s-block consist of two groups: Group1 & Group2.	p-block consist of six group: Group IIIA to Group VIIIA.
They show only +1 & +2 oxidation states	They show a range of oxidation states, +1,+2,+3,+4, +5,+6,+7

Total number of electrons from valence shell configuration = $2+ 2+6+2+3=15$

electronic Configuration of element = $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$.

Period number=3

Group number= 2+3= 6

Q2. Find out the position of an element in the periodic table, which has electronic configuration

Ans. $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$ & In what respect metals differ from metalloids?

Total number of electrons from valence shell configuration= $2+2+6+2+3=15$

Electronic configuration of element $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$.

Valence shell has configuration= $3s^2, 3p^4$.

Period number= 3

Group number= 2+4=6

Hence, S is present in, 3rd period of Group VI

($10+6=$ Group 16).

Metals	Metalloids
They show typical metallic properties.	They show properties of both metal & non-metals
They are present on left hand side of periodic table.	They are present in almost middle of periodic table.
They are present in all block of periodic table i.e s-block, p-block & d & f block	They are present in the p-block of periodic table.

Q3. A cation is smaller than its atom. Justify & Describe the variation of electron affinity in the periodic table.

Ans. It is due to two reasons.

- (i) Due to loss of electrons, effective nuclear charge on remaining electrons increases.
- (ii) Due to loss of electrons the electron-electron repulsion decreases within the valence shell.

Along Period: As you move from left to right in a period, electron affinity increases due to increase in nuclear charge & decrease in atomic size.

In a Group: The electron affinity decreases from top to bottom in group due to increase in atomic size & shielding effect.

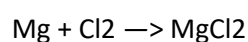
Q4. Explain the reaction of sodium & magnesium with-

Ans. (i) Oxygen (ii) Chlorine (iii) Water & Discuss the variation in nature of oxides of period 3 elements.

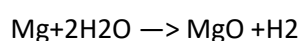
a) Reaction with Oxygen: $2\text{Na} + \text{O}_2 \rightarrow \text{Na}_2\text{O}$ (sodium)



b) Reaction with Chlorine: $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$



c) Reaction with water: $2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2$



In period 3 elements the basic character of oxides decrease along a period. This indicates a decrease in metallic character. It also indicates a decrease in electronegativity difference blw these elements & oxygen.

Q5. Evaluate the significance periodic table & how does the periodic table help in understanding the properties of elements?

Ans. (1).- The modern periodic table arrange elements in order of increasing atomic number.

(2).- The elements arranged in a manner such that their properties repeat in regular manner.

(3).- The table help understand electronic configuration of element & chemical bond.

(4).- The table separate metal, non-metals & metalloids.

- Elements within the same group generally have chemical properties because they have similar electronic configuration.

- The properties of element can be predicted on the bases of their position in periodic table.

Q6. Why is molten sodium chloride a good conductor of electricity? Explain why electrons affinity of fluorine is less than that of chlorine?

Ans. NaCl does not conduct electricity in solid state because it does not have any mobile electrons or free ions to move. However it is a conductor in molten state & in aqueous solution due to ionization.

It is because, F has very small size & seven electrons in 2s & 2p sub-shells. These electrons form a thick electronic cloud- which repel the incoming electrons. So, the electron affinity decreases.

Q7. Why Na₂O is more basic than MgO. Give two reasons? Show the amphoteric nature of Al₂O₃?

Ans. Magnesium oxide is less basic than sodium oxide because the oxide ions are not much free- In sodium oxide case, the solid is held together by attractions between (Na⁺, monovalent) & (O²⁻) ions.

In the magnesium oxide case, the solid is held together attractions are between (Mg²⁺, divalent) & (O²⁻).

Aluminum oxide reacts with hot dilute hydrochloric acid to give aluminum chloride solution.

Al₂O₃ + 6HCl → 2AlCl₃ + 3H₂O shows acidic nature.

Aluminum oxide also shows basic nature.

It reacts with hot concentration sodium hydroxide solution; aluminum oxide reacts to give a colourless solution of sodium tetrahydrox aluminate.

Q8. What is electron affinity?

Ans. It is the amount of energy released when an electron adds up in the valence shell of an isolated atom to a uni-negative gaseous ion.

It is called first electron affinity.

Example: X (g) + e⁻ → X⁻ (g) + electron affinity

Cl (g) + e⁻ → Cl⁻ + 349 kJ/mole

Factors affecting Electron Affinity

Three factors effect electron affinity.

1. Nuclear Charge: Increase in nuclear charge increases the electron affinity and vice versa.

Electron affinity generally increases from left to right in a period. It is because left to right in a period number of protons in the nucleus increases. Thus, the attraction to electrons also increases. So, it becomes easier for the atom to accept an additional electron.

2. Atomic Size: Increase in atomic size decreases the electron affinity and vice versa.

Electron affinity generally decreases from top to bottom in a group. It is because down the group atomic size increases. Larger atoms have electrons farther from nucleus. So, the attraction between the nucleus and the outer electrons is weaker. Therefore, it is more difficult for larger atoms to accept extra electrons.

3. Shielding effect:: Increase in shielding effect decreases the electron affinity and vice versa.

The electron in the inner shells partially shield the outer electrons from the attraction of the nucleus. The more shells there are, the more the outer electrons are shielded. This shielding effect can lead to decrease in electron affinity.

Trends in Periodic Table

Along Period: As you move from left to right in a period, electron affinity increases due to increase in nuclear charge and decrease in atomic size.

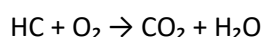
In a Group: The electron affinity decreases from top to bottom in a group due to increase in atomic size and shielding effect.

Chapter-11: Nitrogen and Sulphur

Q1. What gas, omitted from table, makes up most of the remainder of exhaust fumes? & Explain why there is always water vapour in the exhaust fumes?

Ans. The gas that makes up most of the remainder of exhaust fumes is nitrogen (N₂).

The water vapour are always present in the exhaust fumes. Because combustion of hydrocarbons (fuel) in the engine always produce water. The general combustion reaction for hydrocarbons.



Q2. Define Photochemical Smog & formation of PAN.

Ans. In photochemical smog, photo means light & smog is a mixture of smoke & fog.

Photochemical smog is a mixture of the pollutants that are formed when oxides of nitrogen & volatile organic compound (VOCs) react with sun light, creating a brown haze

Formation of PAN: Peroxyacetyl nitrate (PAN) is a highly reactive compound that is a major constituent of photochemical smog. PAN is a secondary pollutant so does not emit directly from the sources but is formed in atmosphere through the reaction described above.

Q3. Differentiate between nitrification and de-nitrification?

Nitrification	De-nitrification
It is an oxidation process. It is a natural process that occurs in the environment where nitrogen compound /ions such as the ammonium(NH_4^+) is oxidised to nitrite ion (NO_2^-) which is further oxidised to nitrate ions (NO_3^-).	It is a biological process where the reduction of the nitrate ion take place in which the nitrogen is released back into atmosphere in their gaseous form nitrogen gas (N_2)

Q4. Uses of sulphur & sulphuric acid?

Ans. Fertilizer: Sulphur compounds like ammonium sulphate are used in the Fertilizers to provide sulphur to crops.

Gun Powder: Gun powder typically consist sulphur , charcoal & potassium nitrate.

Rubber Industry: Sulphur is used in the vulcanization of rubbers.

Uses of Sulphuric acid:

Sulphuric is an extremely important industrial chemical & one of the most widely produced compounds globally. It is highly stable & has a wide range of application in various industries.

Chapter 12: Halogens

Q1. Give the oxides of oxidizing power of halogens? & How does the reactivity of halogens with the hydrogen vary?

Ans. "The elements in group 17 (or Group VII-A) are called halogens"

Halogens are Flourine, Chlorine, bromine, Iodine as taines & oganesson (astatine is radio-active element, oganesson is artificial radio-active element).

Flourine (F_2) → Pale yellow gas

Bromine (Br_2) → Reddish brown liquid

Chlorine (Cl_2) → Greenish yellow gas

Iodine (I_2) → Shiny greyish black solid

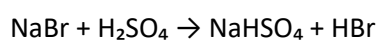
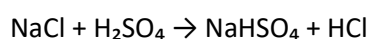
$\text{F}_2 > \text{Cl}_2 > \text{Br}_2 > \text{I}_2$

The bond energy of the halogens decreases with in the group, the atomic size of halogens increases, resulting in weaker intermolecular forces. Hence, the ability of halogen atoms to attract hydrogen atoms is reduced.

Q2. Why is the strength of HF different from HI? & write reaction of Sulphuric acid with NaCl & NaBr?

Ans. "The strength of hydrogen-halogen bond is depended on the electronegativity difference b/w hydrogen & halogen atoms."

A large electronegativity difference results in a stronger bond



Q3. Evaluate the reactivity of hydrogen & halide ions?

Ans. Reactivity of halide ions:

The reducing power of an ion depends on how easily it can lose electrons.

Down the group the reducing power of halide ions increases due to decreases in attraction b/w nucleus & valence electrons

Q4. What is a disproportionation reaction? & Describe the silver nitrate test for the identification?

Ans. When the aqueous solution of silver nitrate is treated with aqueous solution of halide salt, Coloured Precipitates are formed:

- Chloride forms white ppt. remains un-dissolved on adding aqueous ammonia.
- A bromide forms cream off - white ppt, which partially dissolves in aqueous ammonia.
- Iodide forms yellow ppt. which remains un-dissolved in aqueous ammonia.
- Chlorine react with NaOH in two ways. The reaction depends on the temperature & concentration of NaOH.

Chapter-15: Organic Chemistry

Q1. Define catenation & Justify that this property of carbon is responsible for so large number of organic compounds & How was vital force theory rejected by Friedrich wohler?

Ans. The ability of carbon atom to form long chain & ring is called catenation.

Carbon shows catenation because it make strongly carbon-carbon bonds & stable structures of the molecules due to its small size. The presence of multiple bonds, a bonds, attachment of atoms of atoms other than carbon / hydrogen bound carbon / hydrogen isomers vast in nature. compound In 1820, drs Jacob Berzelius said that a spiritual force called vital force is needed to synthesize organic compounds. Therefore, organic compounds cannot be synthesized laboratory. This is vital force theory.

Q2. Differentiate between saturated & un-saturated hydrocarbons. Give three examples for each.

Saturated	Un-Saturated
These have only single covalent bonds between carbon atoms in their molecule.	These have double & triple bonds between carbon atoms in their molecules.
Each carbon atom is bonded to four atoms.	The double bond & triple bonded carbon atom are bonded to less than four atoms.
Carbon atoms have reach their maximum combining capacity no more atom can be attached to carbon atoms.	More atoms can be attached to carbon atoms of double & triple bonds.
They have high hydrogen to carbon ratio.	They have low hydrogen to carbon ratio.

It is symbolic representation of organic compounds which shows simplest whole number ratio b/w elements in them.

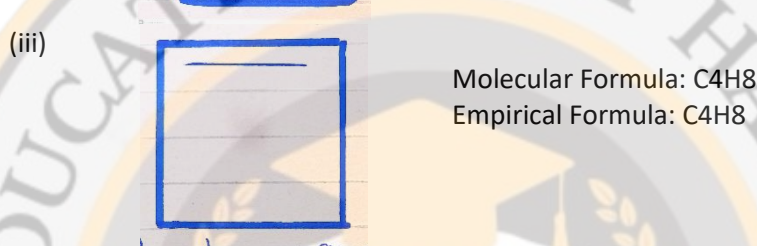
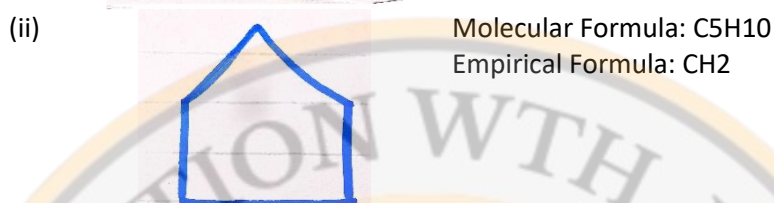
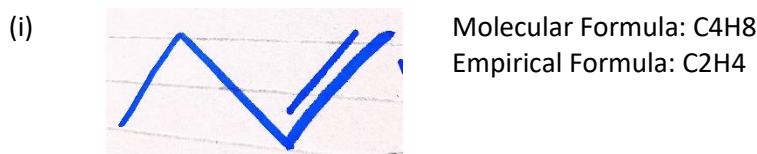
Organic molecules may have may some or different ratio b/w atoms in their molecular formulae.

Example:

Compound	Molecular	Empirical Formula
----------	-----------	-------------------

Ethanol	C ₂ H ₆ O	C ₂ H ₆ O
Ethanoic acid	C ₂ H ₄ O ₂	CH ₂ O

Q3. Deduce the molecular & empirical formula of the following compounds & what is the importance of functional group? Discuss.



An atom or group of atoms which is attached to carbon chains or rings & determines the characteristic properties of organic compounds is called "Functional Group."

Importance:

- Functional groups are the basis for nomenclature (naming) of organic compounds.
- Functional groups help to classify organic compounds into classes/families. All compounds with the same functional group belong to same class.

Q4. Give systematic names to following compounds? What is alkyl group? Name & draw 2D displayed structural formulae of all possible alkyls obtained different isomers of propane & butane?

Ans. (i) CH₃CH₂COCl Propanoyl chloride

(ii) CH₃CN Ethanenitrile

(iii) CH₃OCH₃ Methoxy methane

It is a group obtained by removing any one hydrogen atom from alkanes.

- Alkyls are named by replacing the suffix "ane" of alkane with "yl".
- Thus, the root name of alkane is followed by suffix "yl"

CH₄ → CH₃

methane methyl

Q5. Define electrophile & nucleophile with two examples each? & Define chain isomerism. How can you differentiate b/w metarism & positional isomerism?

Ans. **Electrophile:** An electrophile is an atom or group of atoms which is electron deficient.

Example: Carbocation (CH₃)⁺, H⁺

Nucleophile: A nucleophile is an atom or group of atoms which is rich of electron.

Example: hydroxide ion(OH⁻), Bromide ion(Br⁻)

Chain Isomerism: Chain isomers have same molecular formulae but different structural formulae with respect to length of carbon chain.

H₃C - CH₂ - CH₂ - CH₃

n-Butane

Positional Isomerism: Positional isomers have similar molecular formulae but the different position of functional groups on the principal carbon chain.

Functional Group Isomerism: Functional group isomers have same molecular formula but different functional group.

CH₃ - CH₂ - CH₂ - OH

Metamerism: Metamers have same molecular formula, with respect to chain length on both side of functional group.

H₃C - O - CH₂ - CH₂ - CH₃

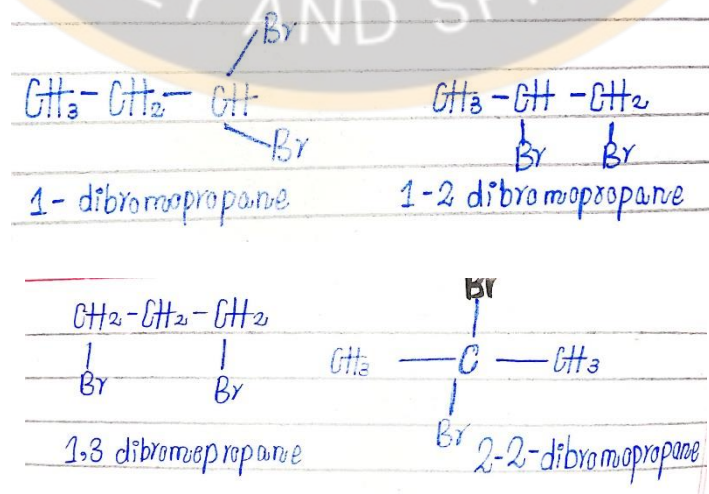
Tautomerism: It is shown by isomers which differs with respect to different position of hydrogen atoms.

(mobile hydrogen):



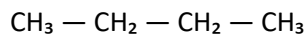
Q6. Draw & name all structural isomers of C₃H₆Br₂. What are aliphatic compounds? Give examples.

Ans. Total isomers are four(4) in number.



Aliphatic organic compounds have straight chain or branched chain molecules. Cyclic molecules, except benzene are also included in aliphatic compounds. They may be both saturated & unsaturated.

Examples:

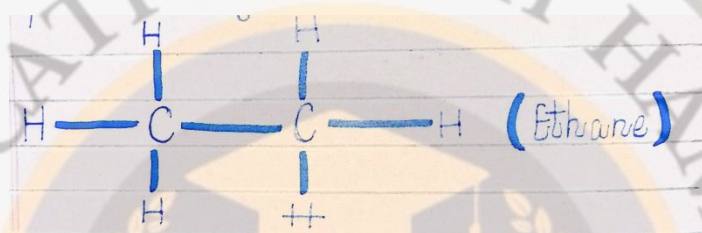


Butane

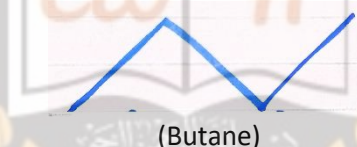
Q7. Define skeleton & 2D displayed structural formulae. Give three examples of each & Derive molecular formula of a compound with the molecular mass 26 amu & empirical formula CH.

Ans. Full Structural Formula:

A full structural formula indicates all atoms & bonds among them in a molecule.



Skeleton Formula: Some organic molecules are very large & complex so it is quite difficult to draw their structures. Skeleton formula is the most easy-to-draw representations for complex organic molecules.



Empirical Formula:

Molecular mass = 26 amu

Empirical formula mass of CH = 13

i.e.

$n = \text{molecular mass} / \text{empirical formula mass}$

$$= 26 / 13 = 2$$

Since,

Molecular formula = n (empirical formula)

Therefore, = $2(\text{CH}) = \text{C}_2\text{H}_2$

Q8. What is homologous series? Write its features & What is free radical substitution reaction of alkanes? Explain homolytic & heterolytic fission. Define nucleophilic substitution of electrophilic addition reaction. Give examples of each.

Ans. Homologous Series: The group of the organic compounds which have same functional group but be different length of carbon chain is called homologous series.

Features of homologous Series:

- 1. Each homologous series has its own functional group.
- 2. A homologous series has gradual change in their physical properties with changing the size of carbon chain.
- 3. They have similar method of preparations.

Electrophilic Addition Reaction: This type of reaction involves the attack of an electrophile on the multiple bond

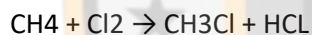
Nucleophilic Substitution Reaction: In these reactions, a nucleophile replaces an atom or group of atoms in a molecule.

Nucleophilic Addition Reaction: In this case, a nucleophile is added to a molecule, usually aldehydes & ketones.

Q9. What is free radical substitution reaction of alkanes? Explain bromination of methane with mechanism. Explain the difference b/w homolytic & heterolytic fissions of the covalent bonds. How can homolytic fissions initiate the halogenation reaction of alkanes?

Ans. Free Radical Substitution Reaction of Alkane:

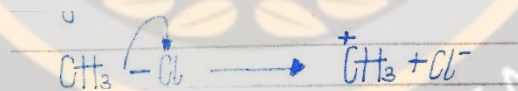
- In such chemical reactions, a free radical is formed first.
- The free radical then replaces hydrogen atoms of alkanes one after the other.



Homolytic Fission: "It involves equal splitting of covalent bond between two atom in a molecule."



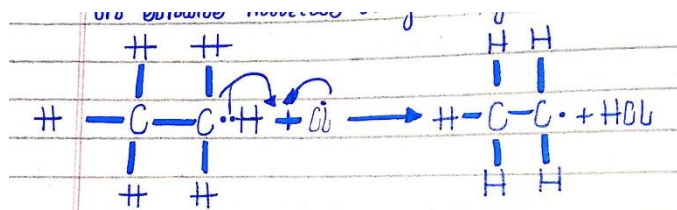
Heterolytic Fission: It is unequal breakage of covalent bond between atoms in organic molecules.



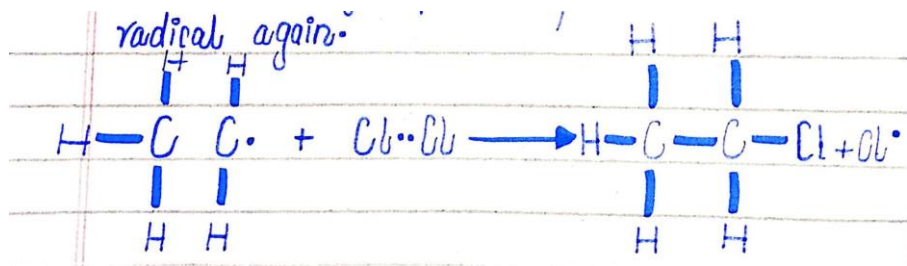
Chapter-16: Hydrocarbons

Q1. Justify why the second step in halogenation of alkanes is named propagation? sp Draw the structure of ethane molecule?

Ans. The second step of free radical mechanism is named propagation because it occurs like a chain reaction. First chlorine free radical attacks on ethane molecule to give ethyl free radical.

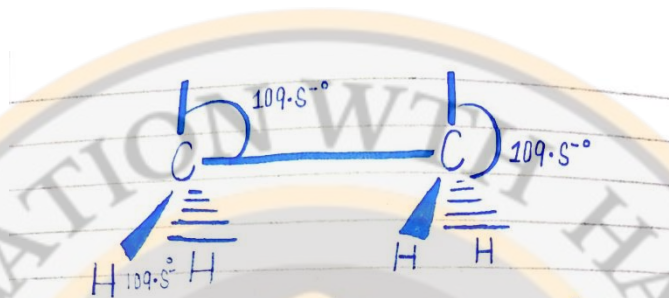


Then ethyl free radical attacks on chlorine molecule to give product & chlorine free radical again.

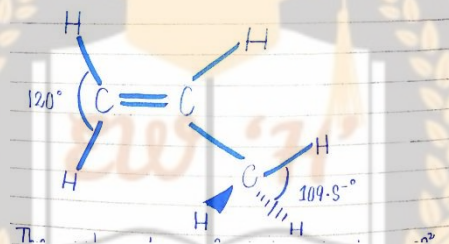


This steps are repeated again & again like a chain reaction. The products are formed in this step.

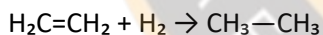
Structure of Ethane:



Q2. Draw stereo chemical formula of propane & Alkenes shows addition reaction while alkanes cannot. Why?



The carbon atoms of double bond are sp² hybridized so they are planar with bond angles of 120°. Alkenes show addition reaction in which a molecule is added to alkenes across double bond. Such reaction involves the conversion of double covalent bond to single covalent bond.



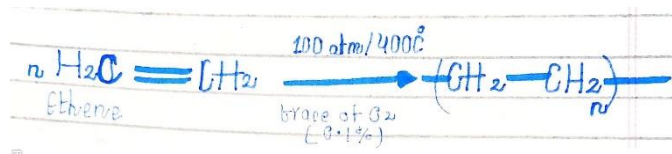
Q3 Halogen is a non-polar molecule. How can it act as electrophile in halogenation of alkenes & How can we prepare poly(ethene) by polymerization etc?

Ans. When a non-polar bromine comes close to ethene molecule in ethene molecule due to repulsion b/w electron cloud of bromine & ethene molecule, the halogen molecule gets polarized.

Polymerization: In it, small molecule called monomers react together & form large molecules called polymers.

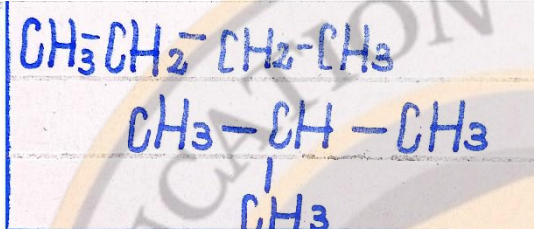
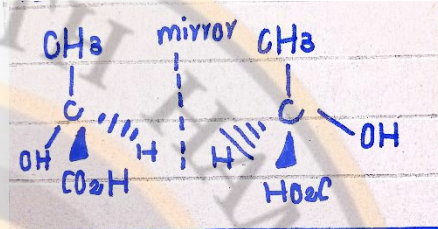
- The word polymers comes from the Greek language "polymeros" - 'poly' means many, 'meros' means parts.

Ethene molecule can polymerize at 400°C temperature & 100 atm pressure in the presence of traces of oxygen from the polyethylene polymised



Ethene gives harder & high melting poly (ethene) polymer in presence of mixture of triethylaluminium, $\text{Al}(\text{C}_2\text{H}_5)_3$ & titanium tetra chloride, TiCl_4 , which acts gets as catalyst (Ziegler-Natta catalyst)

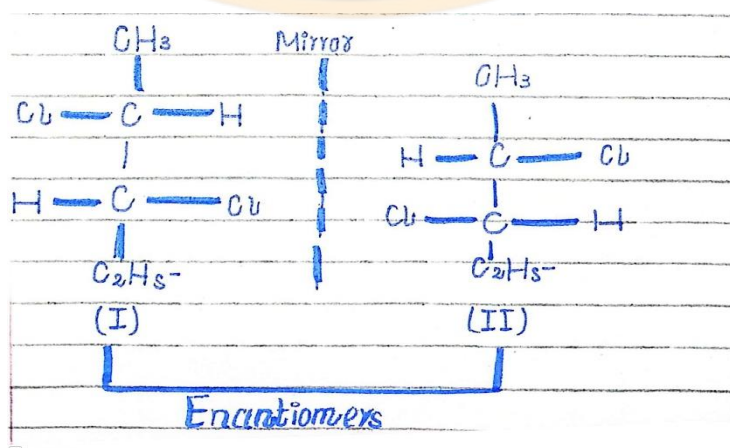
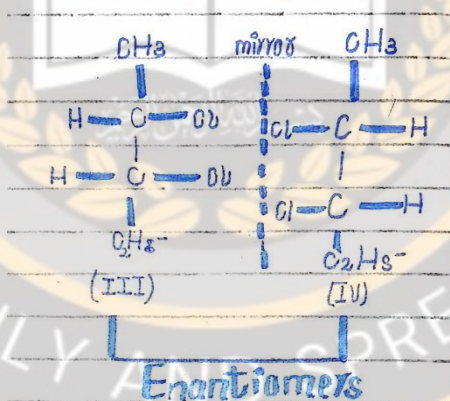
Q4. What is the difference b/w structural & stereo isomerism? Support your answer with 2D & 3D formulae & How can we differentiate enantiomers & diastereomers?

Structural isomerism	Stereoisomerism
The structural isomers have same molecular formulae but different arrangement of the atoms in molecules.	The stereoisomers have different arrangement of atoms in the space
The interconnectivity of atoms may change.	The interconnectivity of atoms is same
It is 2D dimension.	It is 3D dimension
	

Enantiomers are non-superimposable mirror images of each other.

Enantiomers have same physical & chemical properties except optical activity & interaction with other optically active compounds. They are always optically active.

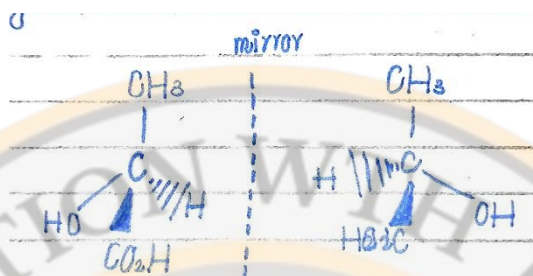
Examples: consider isomers of 2-3 dichloropentane



Q5. All chiral molecules are not necessarily optically active but all asymmetric molecules are optically active. Justify this statement by an example & Generally, hydrocarbons are less reactive. Arrange alkanes, alkenes, & alkynes in their reactivity order & justify?

Ans. The chiral carbon has four different groups attached to it. Compounds containing chiral carbons are generally optically active.

e.g lactic acid

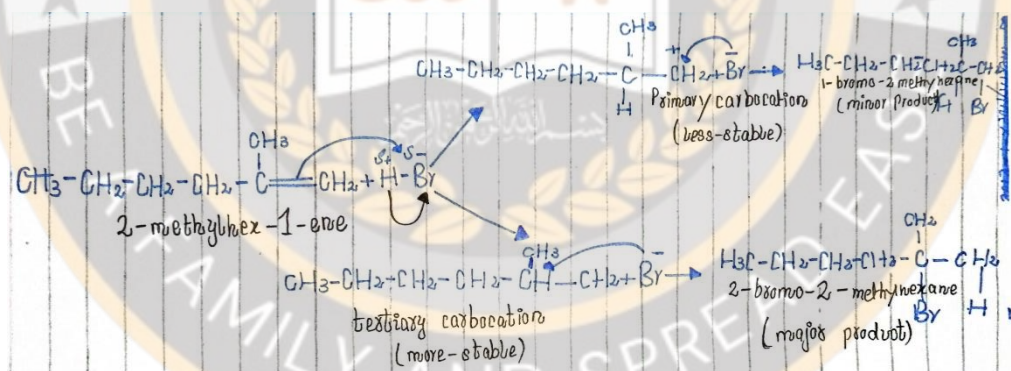


In alkanes, only sigma bond is present, which is difficult to break. These are also non-polar. Hence, alkanes are least reactive.

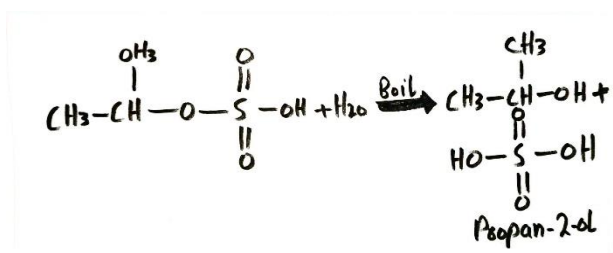
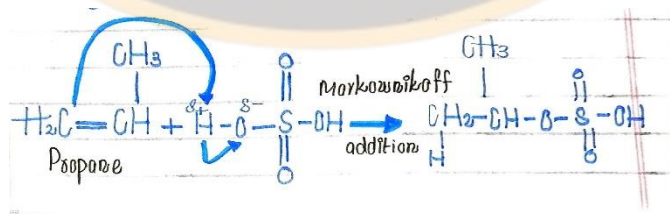
Q6. Define Markovni rule- Give mechanism of hydrobromination of 2-methylbut-2-ene? & Hydration of alkenes is acid-catalyzed reactions . Justify & exemplify this statement by propene?

Ans. It states: When an unsymmetrical reagent (H-Br) attacks on un-symmetric alkene (propene) > the hydrogen atom bonds itself to that double bond carbon which has more hydrogen atoms.

Structure

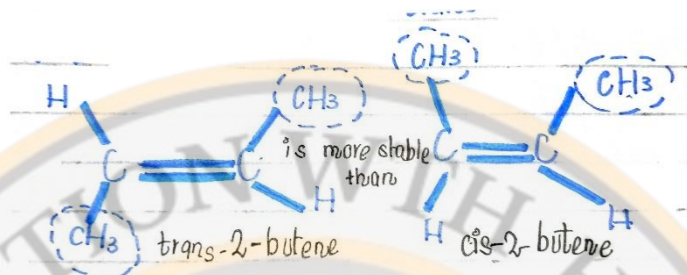


Hydration of alkenes is an addition reaction which occurs in the presence of a mineral acid e.g H₂SO₄

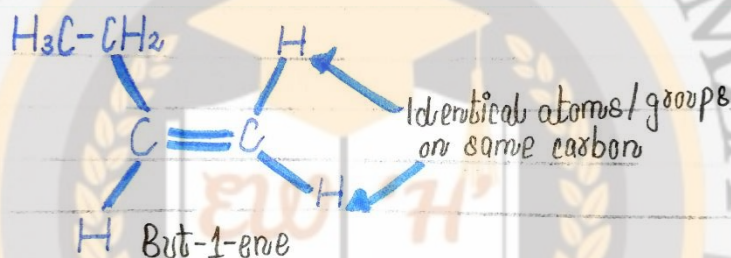


Q7. Explain why cis-alkenes are generally more stable than their trans-isomers? & Why there are no geometric isomers of but-1-ene?

Ans. The trans-alkene is generally more stable than cis-alkene. It is because when R-groups are on opposite sides of the double bond, there is less steric interactions & hence more stability e.g trans-2-butene > cis-2-butene



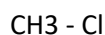
In but-1-ene no geometrical isomers are possible. It is because one of double bonded carbons has two identical groups attached to it.



Chapter- 17: Halogenoalkane

Q1. What are primary, secondary & tertiary alkyl halides? & what are nucleophilic substitution reactions?

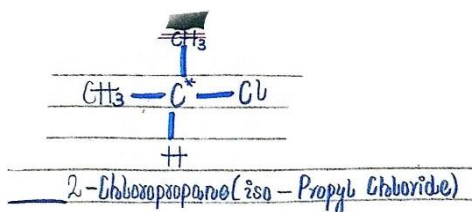
Ans. Primary Alkyl Halides: Alkyl halide in which halogen atom is attached with a primary carbon.



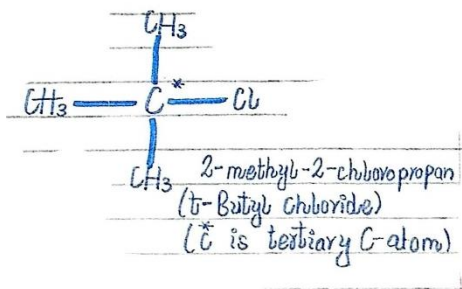
methyl chloride

Chloro methane

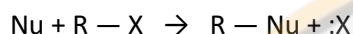
Secondary Alkyl Halides: Alkyl halide in which halogen atom is attached with a secondary carbon atom.



Tertiary Alkyl Halides: Alkyl halid in which halogens atom is attached to a terbiary carbon.



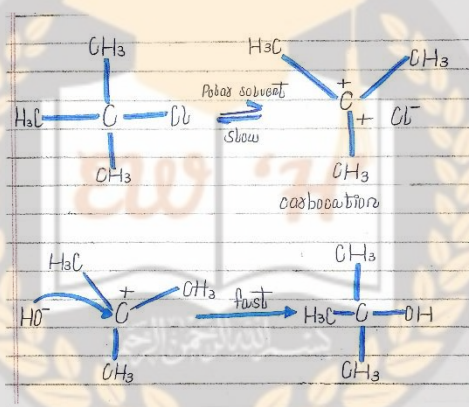
The reaction where halogen of alkyl halide is substituted or replaced by attacking nucleophile is known as nucleophilic substitution (SN) reactions.



attacking nucleophile leaving Nucleophile

Q2. Tertiary alkyl halides show SN1 reaction mostly, why? & what are elimination reaction?

Ans. Hence, in step I, the tertiary alkyl halide are ionized to give carbocation which is planar. Then in step II, the nucleophile attack to give products. This two-step mechanism is SN1 mechanism.



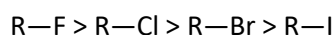
The chemical reaction in which two groups (halogen & B - hydrogen form alkyl halide) are eliminated from two adjacent carbon atoms to form carbon double bond is called elimination reaction.

Q3. Which factors decides the reactivity of alkyl halides? & what are the diazonium salts?

Ans. There are two main factors which control the reactivity of halogen alkanes:

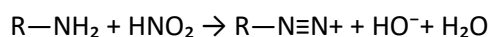
1. Bond polarity of C-X bond.
2. Bond energy of C-X bond.

On the basis bond polarity reactivity of alkyl halides decreases in following order.

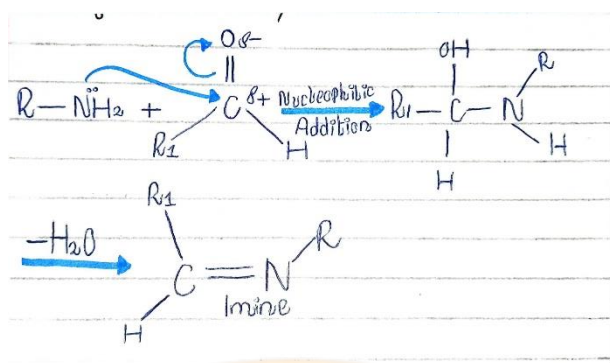


However, actually reactivity of alkyl halides depends upon the strength of C-X bond.

When amines react with nitrous acid, diazonium compound are formed.



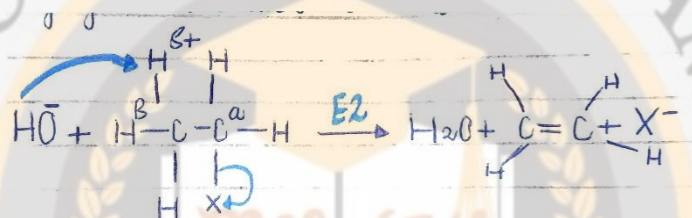
Q4. How can nucleophilic addition of a primary imine? & Amines are more basic than analogous alcohols why?



Amines may act as bases towards acids. They are more basic than alcohols & ethers. It is because they have single electron pair on nitrogen atom which is readily donated. (In alcohols) there are two lone pairs of electron on oxygen atom.

Q5. What is the difference between a base & nucleophile? & What is a poor leaving group?

Ans. When a specie carrying negative charge or a lone pair of electron attacks on the hydrogen atom then it is called a 'base'.



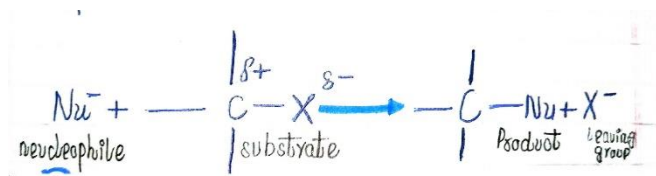
If such specie attacks on carbon atom then it called nucleophile.



A poor leaving group is that which is not easily removed desing substitution or dimited reaction. Generally, the groups which become stable independent species after leaving are good leaving groups vice versa.

Q6. Why alkyl halides undergo nucleophilic substitution reaction? Exp Also define between Bond polarity & Bond Energy ?

Ans. Alkyl halides have electrophilic carbon atom . Thus, nucleophile can attack an electrophilic carbon & halide ion is removed as leaving group. Thus, alkyl halides give nucleophilic substitution reaction.



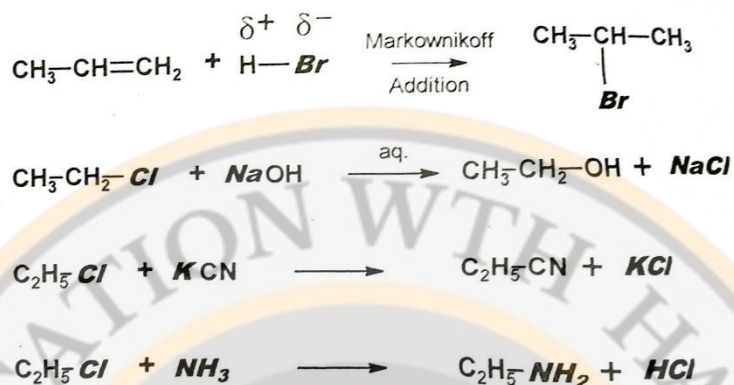
Bond Polarity : The molecule of alkyl halide is polarized due to the greater electronegativity of halogen then C.

Hence, carbon gets partial positive charge while halogen gets partial negative charge

Bond Energy : Energy required to break the bond of same type in one mole bond is known as bond energy.

The greater the bond energy, the more energy will be required to break bond.

Q7. Complete the following reaction.



Chapter 18: Alcohol

Q1. What are alcohols? How are they classified?

Ans. So Give the mechanism of dehydration of alcohol to form ethers?

"The aliphatic organic compounds containing hydroxyl group -OH as functional group are called alcohols"

Examples: CH₃ - CH₂ - OH

Monohydric alcohol

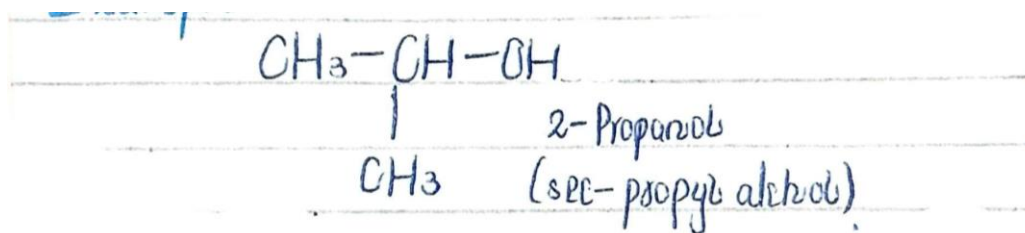
(1). Primary alcohol: In these, carbon atom attached to OH group, is directly attached to one or no carbon atom.

Examples: CH₃ - OH

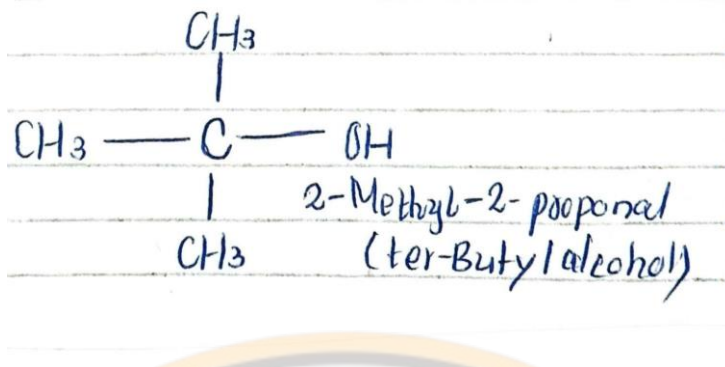
methanol

(2). Secondary alcohol: In these, carbon atom attached to OH group, is directly attached to two carbon atoms.

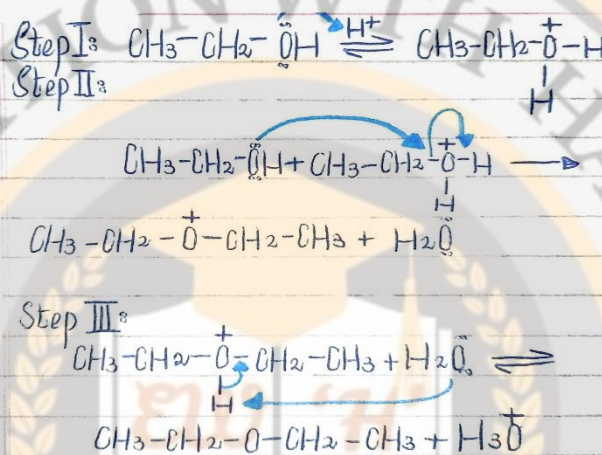
Example:



(3). Tertiary alcohol: In these, carbon atom attached to OH group is directly attached to three carbon atoms



Q2.

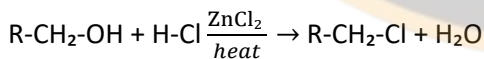


Q3. Why an alcohol is less acidic than water?

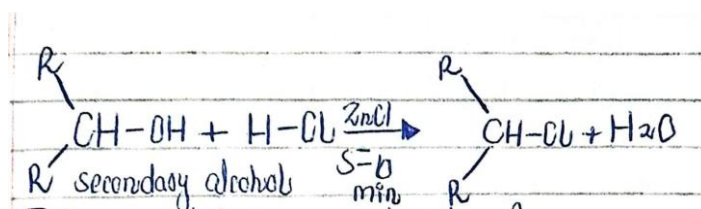
Ans. Alcohols & water both contain -OH group. The -OH group in alcohols is less polar than -OH group in water. The electron releasing nature of alkyl group decreases the polarity of -OH bond. This makes alcohols weaker acid than water

Q4. How will you distinguish b/w primary, secondary & tertiary alcohols? Explain with reaction? & Give at least two methods for the preparation of ethanol?

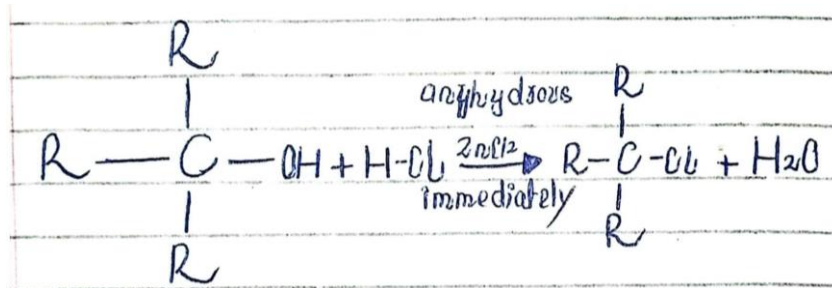
Ans. (i) Primary alcohols forms an insoluble prim-alkyl chloride on heating.



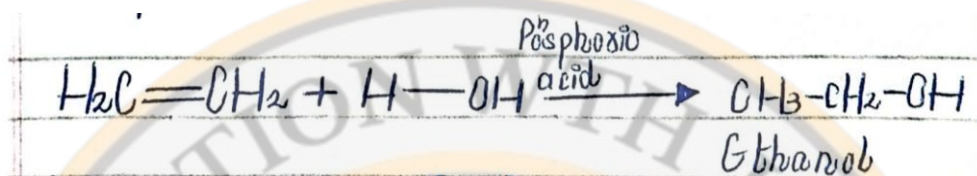
(ii) Secondary alcohol forms an insoluble sec-alkyl chloride in 5-10 minutes



(iii) Tertiary alcohol immediately insoluble bayes of a ter-alkyl choside



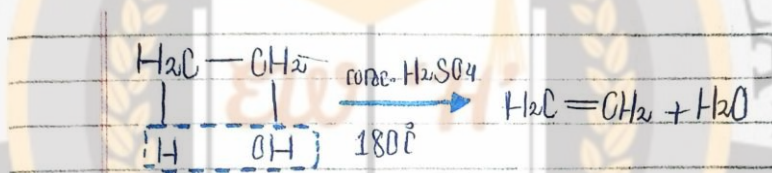
(c) Alkenes react with steam in presence of the phosphoric acid as catalyst at 300°C & 60-70 atm. pressure to form alcohol-



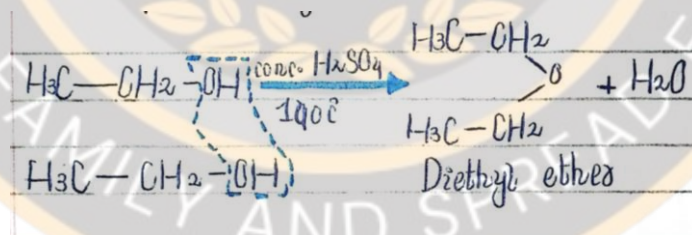
Q3. : Ethanol gives different products with conc. H2SO4 under different conditions & tertiary alcohols are not oxidized by acidified KMnO4 or K2Cr2O7?

Ans. Ethanol react with conc. H2SO4 & give the different products at different temperatures

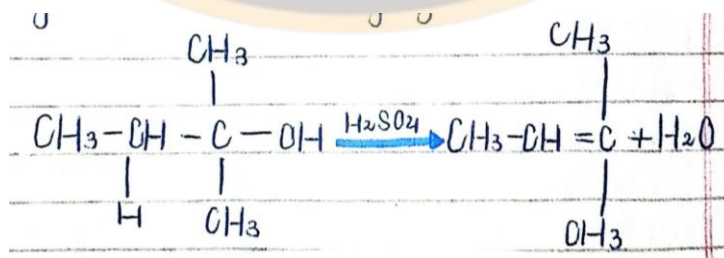
(i) high temperature, low alcohol conc.:



(ii) low temperature, high alcohol conc.:

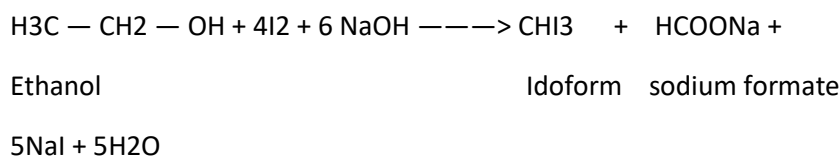


Tertiary alcohols are not oxidized because they donot contain α - hydrogen.

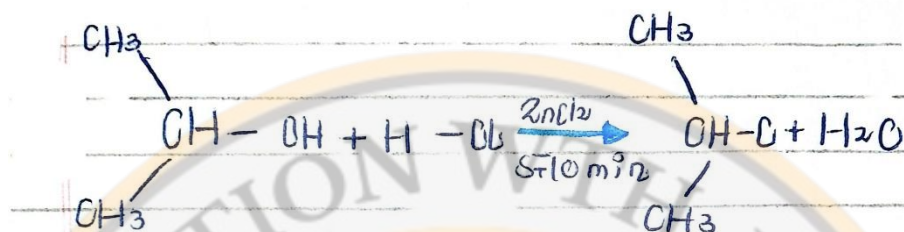


Q4. How will you distinguish between?

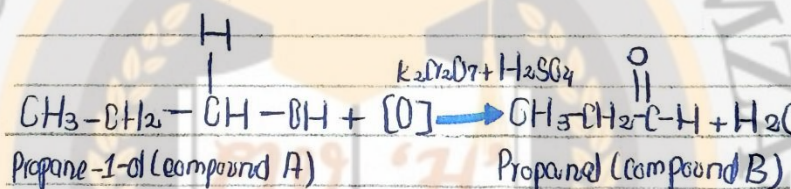
(i) Methanol & ethanol: Ethanol -give iodoform test. Yellow ppt of iodoform are formed when ethanol is treated NaOH/I2.



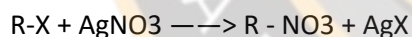
(1) 1-Propanol & 2-propanol. 2-propanol is a secondary alcohol. When it reacts with Lucas reagent (HCl+ZnCl₂), it produces oily layer of insoluble sec-alkyl halide within 5-10 minutes



Q5. An organic compound A has molecular formula C₃H₆O which on oxidation forms a compound B which on oxidation form another compound C. Identify structural formulae of compounds A, B, & D. if How will you distinguish b/w the following?

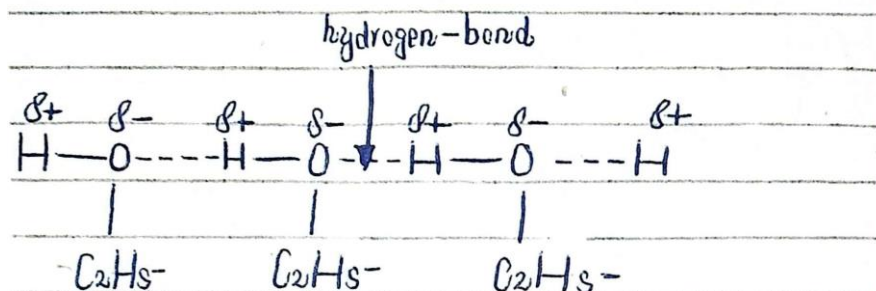


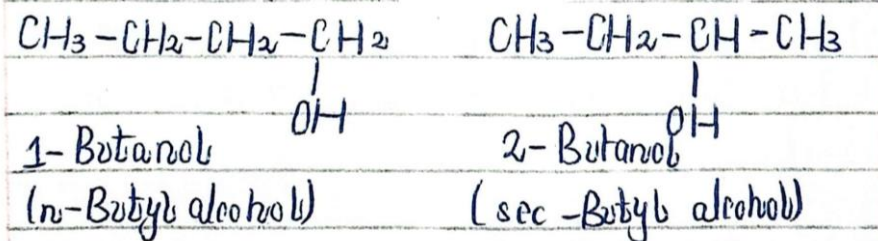
(i) An alcohol & an alkyl halide: When an alkyl halides is heated with AgNO₃ in alcohol, the halogen atom is replaced by NO₃ group to generate a nitroalkane. This reaction can be used to find which halogen is present in alkyl halide



Q6. Explain why the boiling points of alcohols are much higher than those of alkanes & ethers? How many isomeric alcohols are possible for C₄H₉OH? Draw their structural formulas & write both common & IUPAC names?

Ans. Alcohols & phenols have strong hydrogen bonding, while ethers have weak van der Waals forces among their molecules e.g H-bonding is represented as.



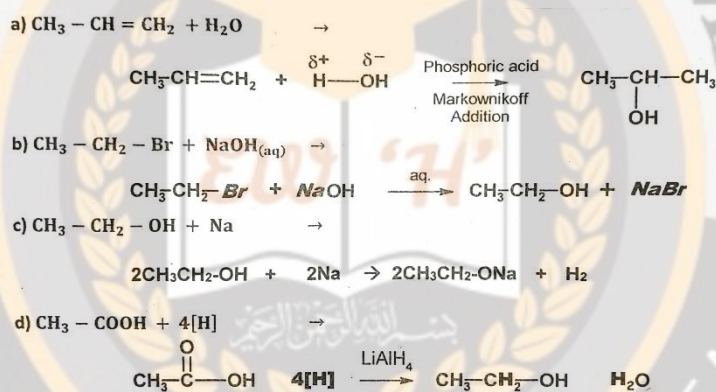


Q7. Why tertiary alcohol cannot be oxidized? & Compare the boiling point of water, ethanol & diethyl ether?

Ans. Tertiary alcohols are not oxidized because they do not contain hydrogen.

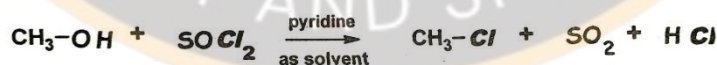
In water, two hydrogen bonds are formed per molecule while in ethanol only one hydrogen bond is formed per molecule. Thus, due to extensive hydrogen bonding in H₂O, its boiling point is greater than ethanol.

Q8. Complete the following reaction.



Q9. How can you convert it?

(a) Methanol into chloromethane



(b) Ethanol into bromoethane



Chapter-19: Carbonyl Compound

Q1. How can you distinguish ethanol from methanol? & Describe briefly the nucleophilic addition mechanism of carbonyl compound?

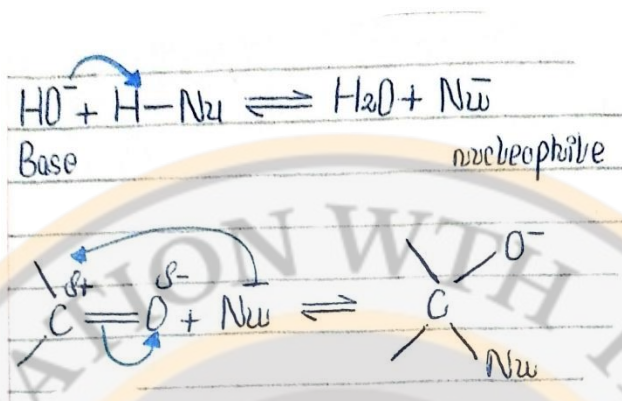
Ans. Ethanol gives iodoform with iodine in the presence of NaOH. Formation of yellow crystals of iodoform indicates ethanol. Methanol does not give iodoform test.



A base catalyzed nucleophilic addition reaction takes place with a strong nucleophilic reagent.

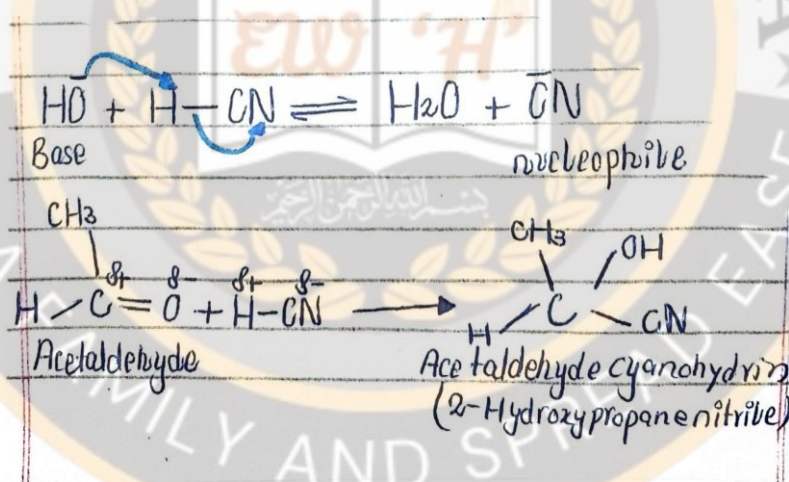
The base reacts with the reagent & generates the nucleophile.

General Mechanism:



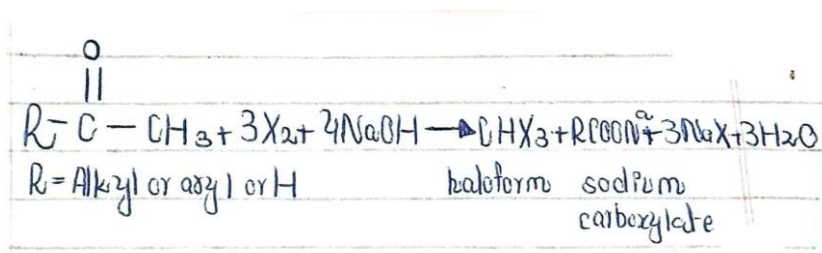
Q2. What is mechanism of HCN addition to the carbonyl compounds ? & How is ethanal converted to lactic acid?

Ans. The base (OH^-) generates cyanide ions which act as nucleophile. This nucleophile reacts with carbonyl compound.

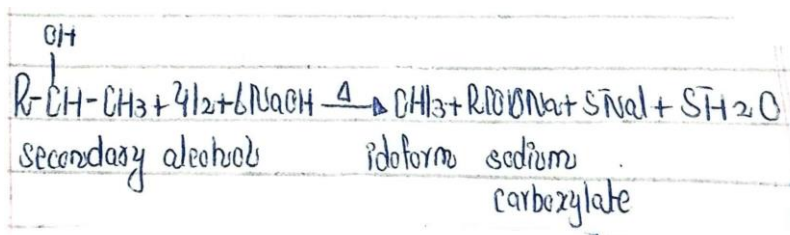


Q3. What is Haloform reaction? which type of the alcohols undergo iodoform reaction?

Ans. Acetaldehyde, ethanol, methyl ketones & secondary alcohols with methyl group at carbon, react with halogens ($\text{Cl}_2/\text{Br}_2/\text{I}_2$) in the presence of sodium hydroxide to give haloform (chloroform, bromoform or iodoform) & sodium salt of acid. This reaction is called haloform reaction.



Secondary alcohols containing the methyl group also undergo this reaction.



Q4. What are aliphatic & aromatic carboxylic acids & How are esters produced?

Ans. Carboxylic acids are generally two types.

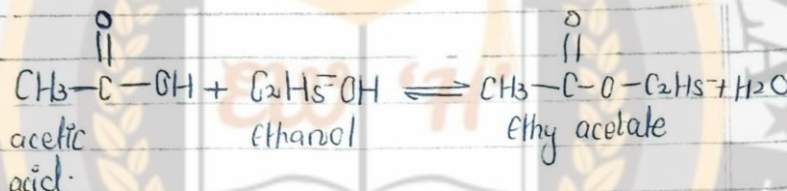
Aliphatic carboxylic acids & aromatic carboxylic acids

The general formula of aliphatic acid is

R - C - OH where R=H or alkyl group



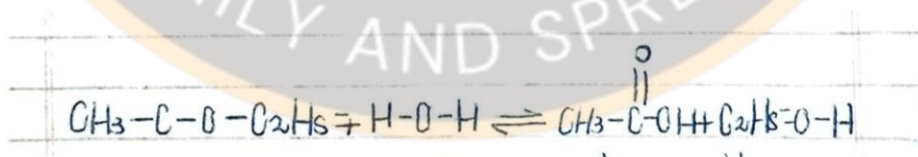
e-g Acetic acid (CH₃COOH).



Q5. Give mechanism for reaction for acid catalyzed esterification? sp what is reactivity of carbonyl group?

Ans. Esters can be hydrolysed to acid & alcohol with dilute acid or dilute alkali & heat.

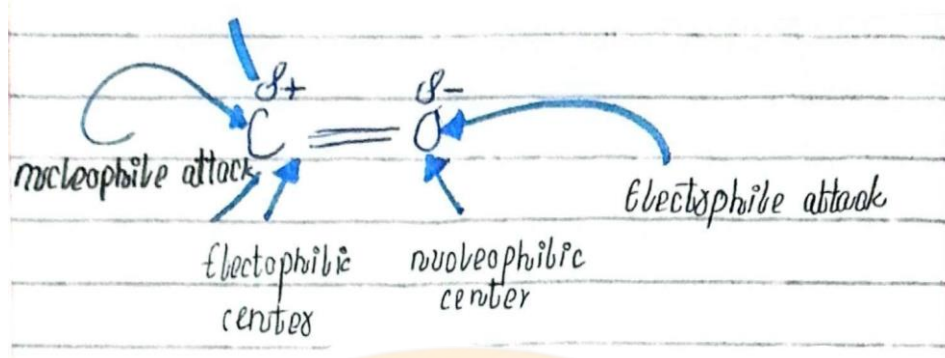
Basic hydrolysis of esters is known as saponification because it leads to formations of alkali salt of carboxylic acid which is soap.



In carbonyl groups there is present a sigma bond & pi bond. Thus, it can give addition reaction.

Due to more electronegativity of oxygen, it attracts shared electrons towards itself.

Thus, it gets e- sp becomes nucleophilic while carbon get s+ charge & becomes electrophilic. Thus, carbonyl group becomes polar & nucleophile or electrophile can attract on it.



Q6. What is minimum numbers of carbons atoms a ketone may contain? Comment on it & write structural of all carbonyl compound having molecular formula C₃H₆O.

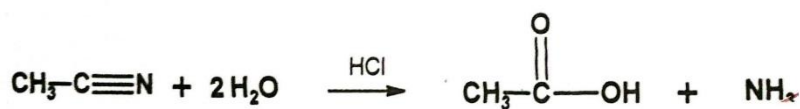
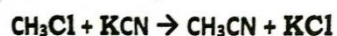
Ans. (i) A ketone is compound in which carbonyl group (C=O) has carbon atom on both sides. Thus, minimum three carbon atoms should be present for ketone. The simplest ketone (CH₃COCH₃) is acetone with three carbons atoms.

Classify each compound according to the functional group?

The two are functional group isomers. One is ketone & other is aldehyde.

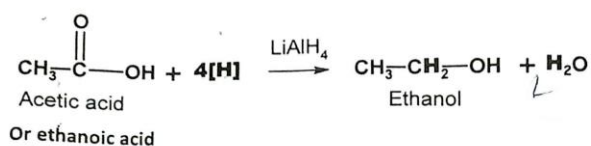


Q7. How can you convert CH₃Cl into CH₃COOH?

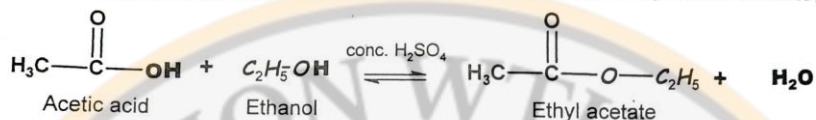


Q8. How can you convert ?

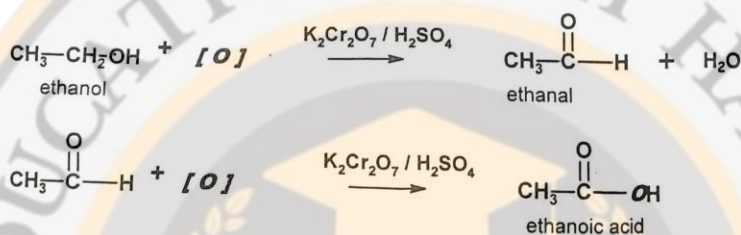
(1). Ethanoic acid into ethanol. Reduction of ethanoic acid with LiAlH_4 gives ethanol



2. Ethanoic acid into ethyl acetate. Ethanoic acid reacts with ethanol in the presence of acid to give ester.



3. Ethanol into ethanoic acid



Chapter-20: Nitrogen compound Amines

Q1. What are amines? Explain amines are more basic than ammonia?

Ans. The group of organic compound containing nitrogen linked with alkyl or aryl groups with a lone pair on nitrogen atom are called amines.

The amines lone pair of electron is present on nitrogen atom. Thus, amines are basic & are nucleophiles because they can donate non-bonding pair of electrons to an electron-deficient center.

Example:



Q2. Amines have high melting/ boiling points as compared to other analogous alkanes? why? Cp methylamine is stronger base than ammonia? Explain.

Ans. The N-H bond is polar bond. It is due to electronegativity difference of two atoms.

Due to polarity hydrogen bonds are formed with other amine molecules.

In amines lone pair of electrons is present on nitrogen atom. Thus, amines are basic & nucleophiles because they can donate non-bonding pair of electrons to an electron-deficient center.

Examples:

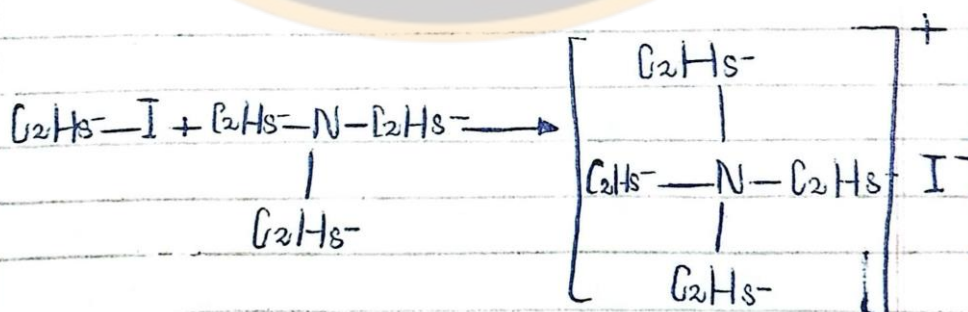
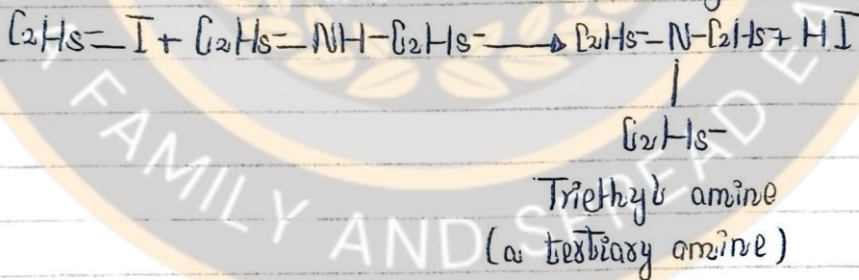
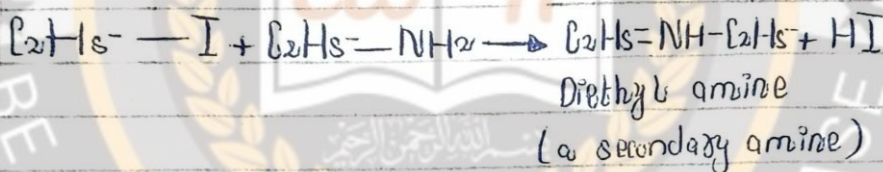
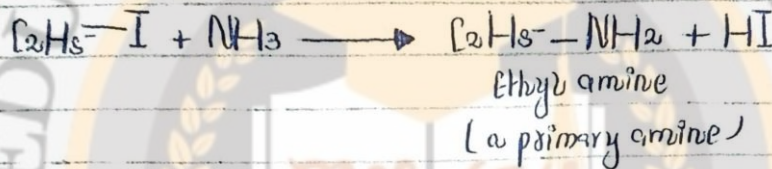


Q3. How would you prepare Diazonium salt? & Discuss the methods of preparation on amines?

Ans. When amines react with nitrous acid in the presence of hydrochloric acid, diazonium compounds are formed. The reaction mixture must be kept below 10°C.



When an alcoholic solution of ammonia is heated with an alkyl halide in a sealed tube at 100°C a mixture of primary, secondary amines & a quaternary ammonium salt is obtained



Salts of amines are produced in acidic medium. Thus addition of strong alkali e.g KOH liberates free amines from their salts. The quaternary salt is unaffected.

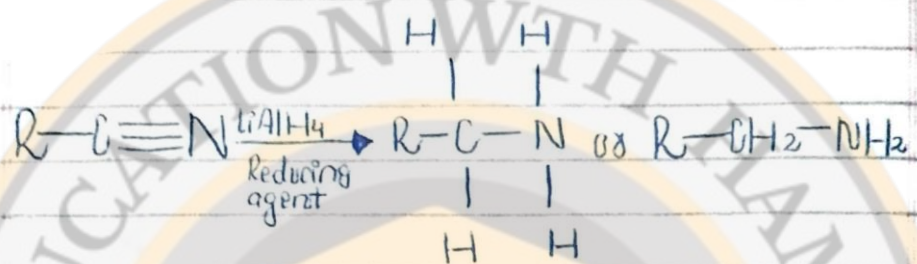
The three types of amines are separated by fractional distillation.

Does alkylation can be avoided by using excess of ammonia but the yield is low.

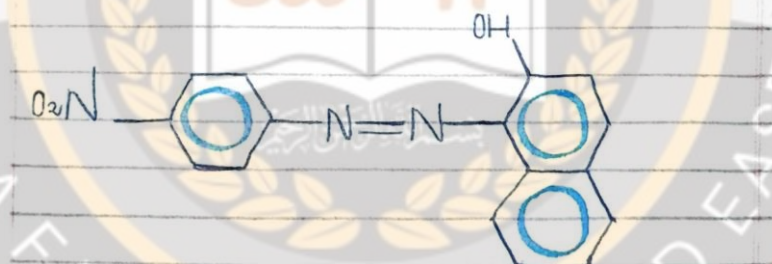
Q4. What is nitriles? How they can be reduced? Evaluate the importance of diazonium salts in everyday life?

Ans. The compounds which contain cyano group (CN) are called nitriles- reduction of alkyl or aryl nitriles gives primary amines.

The reduction may be brought about by reducing agent i.e LiAlH₄, or sodium in ethanol.

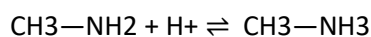


Diazonium salt combined different phenols to give azo dyes. e.g with phenol gives an orange-red dye. Phenyl azo compounds are used as dyes & pigments in textile, cosmetics, paints, & leather industries. e.g para red dye is an azo dye.



Q5. Why aliphatic amines are more basic while aromatic amines are less basic than ammonia. & why secondary amines are more basic than tertiary in aqueous solutions?

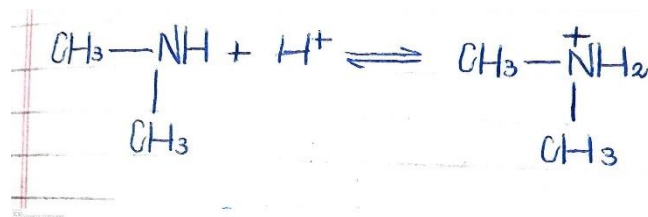
Ans. In aliphatic amines e.g CH_3NH_2 , the lone pair of electron is readily available on nitrogen atom. Hence aliphatic behave as relatively stronger bases.



The basicity depend upon

- (i) the availability of electron pair on nitrogen atom.
- (ii) the stability of corresponding cation.

The cation formed by secondary amine is



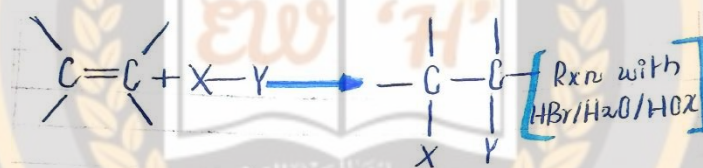
Chapter-21: Organic Synthesis

Q1. Compare & contrast b/w synthesis & retro synthesis & Also differentiate b/w Electrophilic Addition & oxidation reaction?

Synthesis	Retro synthesis
A chemical synthesis for the preparation of organic compounds	A technique to find a suitable route to the synthesize a compound.
It is the process of combining simple rxn to form an organic compound.	It is the process of the working backward from the target organic compound to suggest a suitable route of synthesis

Electrophilic Addition Reaction:

- In these reaction one of two bonds in carbon - carbon double bond is broken & a new single bond is formed from each of two carbon atom.
- In an addition reaction a single product is always formed



Oxidation Reactions

Diol formation or Hydroxylation:

Alkenes can be oxidized to give diols in the presence of cold dilute alkaline solution of KMnO_4 .



BEST OF LUCK