

OROMIA EDUCATION BUREAU



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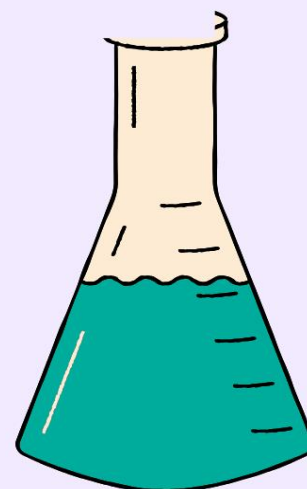
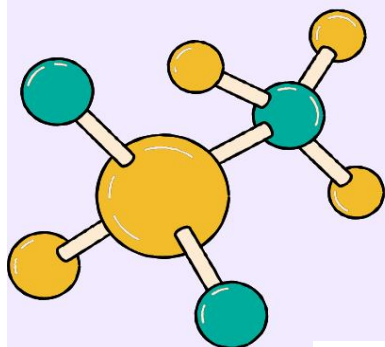
CHEMISTRY

INSIGHT MODULE-I

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Content interface

This Chemistry module has been prepared based on the new curriculum revised by the Ethiopian Ministry of Education.

Its main purpose is to assist Grade 12 students in preparing for the national exam.

- *The module covers most of the Grade 9-12 contents, integrating similar topics to enhance comprehension.*
- *It is structured in a question-and-explanation format, making it easier for students to*
- *It is evaluated by Adama science and technology university stakeholders*
- *It consists of three main sections*
 1. *Introduction*
 2. *Questions with their explanations*
 3. *Self-practice question*

Acknowledgment

First and foremost, I give my deepest gratitude to **Almighty God** for His guidance, wisdom, and strength throughout the development of this module. Without His blessings, this work would not have been possible. I would also like to extend my sincere appreciation to **Mr. Garedew Jima** and **Mr. Mekonnen Kebede** for their invaluable support in developing this module. Their guidance and contributions have played a crucial role in its completion.

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UNIT-ONE

1. ATOMIC THEORY

Unit Out Comes

- ❖ *At the end of this unit, students will be able to:*
 - ❖ *Explain the experimental observations and inferences made by some famous scientists to characterize the atom*
 - ❖ *State the experimental investigation of proton, electron, nucleus and Neutron*
 - ❖ *Apply Aufbau principle, Pauli Exclusion Principle and hund's rule on electron configuration*
 - ❖ *State the atomic theory thought at different*
 - ❖ *Compute the calculations involving atomic structure*
 - ❖ *State the postulates of Bohr model*
 - ❖ *Compare and contrast the thought of Bohr model and Quantum mechanical model*
 - ❖ *Compare and contrast different electromagnetic radiation in terms of wave length, frequency, wave number and energy.*
 - ❖ *Describe the quantum mechanical model of the atom and the related postulates and principles*
 - ❖ *Demonstrate periodic law and how electronic configurations of atoms are related to the orbital diagrams and can explain periodic trends*
 - ❖ *Developing skills on problem solving about Atomic theory regarding to different equations.*
 - ❖ *Demonstrate different Quantitative parameters like wave length, frequency, wave number and energy about atomic theory*
 - ❖ *Describe the four quantum number (n, l, ml, ms) of quantum mechanical model.*
 - ❖ *Explain key concepts about, isotopes, photo electric effect, dual nature of electron, de Broglie wave length, atomic models, electron configuration*

1.1. INTRODUCTION

The smallest piece of an element that maintains the identity of that element is called an atom.

Individual atoms are extremely small. It would take about fifty million atoms in a row to make a line that is 1 cm long. Atoms are so small that it is difficult to believe that all matter is made from atoms.

The concept of Atomic theory started from the ancient Greek philosopher, Democritus

Democritus → Aristotle → **Dalton Atomic theory** → **Modern Atomic Theory**

1.1.1. Atomic models

1. Dalton's atomic model (billiard ball model)

- * It was put forward by the English physicist and chemist John Dalton.
- * It stated that matter is made up of small indivisible particles called 'atoms'.
- * According to this theory, all substances are made up of atoms, which are indivisible and indestructible.
- * While atoms of same element have same size and mass, atoms of different elements have different size and mass.

2. Thomson's model of atom (Plum pudding model, chocolate chip ice-cream model.)

- * It was put forward by J.J. Thomson.
- * It was also called plum pudding model.
- * This model explained the inner structure of atom theoretically.
- * According to Thomson, an atom resembles a sphere of positive charges with electrons distributed inside it.
- * Here, positive and negative charges are equal in magnitude and therefore an atom is electrically neutral.

3. Rutherford atomic model (Nuclear model)

- * It was put forward by the British scientist Ernest Rutherford.
- * The model by Thomson was not able to explain certain results associated with the atomic structure of elements, so this model came into existence.
- * This model described atom as a tiny, dense, positively charged core called nucleus, in which near all the mass is concentrated, around which the negatively charged electrons circulate like planets moving around the sun

4. Bohr's Model (Planetary model)

- * It was put forward by the Danish physicist Neil Bohr.
- * It is similar to Rutherford model.
- * The Bohr model is a planetary model in which negatively charged particles orbit positively charged nucleus similar to planets orbiting the sun.
- * Coulomb force exists between these positive and negative particles.

The Bohr Model of the atom, proposed by Niels Bohr in 1913, is based on quantized energy levels for electrons in atoms. The main postulates of Bohr's model are:

1. Electrons move in fixed orbits (stationary states)

Electrons revolve around the nucleus in specific, discrete circular orbits without emitting energy.

These orbits are called *energy levels* **or** *shells* (designated as $n=1,2,3,\dots$)

2. Quantization of Angular Momentum

*The angular momentum of an electron in an orbit is **quantized** and given by the equation:*

$$mvr = \frac{nh}{2\pi}$$

where:

- m = mass of the electron
- v = velocity of the electron
- r = radius of the orbit
- n = principal quantum number (1, 2, 3, ...)
- h = Planck's constant

3. Energy is emitted or absorbed only during transitions

Electrons do not lose energy while moving in a stationary orbit. However, they can jump from one energy level to another by absorbing or emitting a quantum of energy (photon), given by:

$\Delta E = h\nu$, where: ΔE = energy difference between two levels

h = Planck's constant

ν = frequency of radiation absorbed/emitted

4. Electron orbits correspond to fix energy levels

Each orbit has a fixed energy, and electrons in a particular orbit have specific, quantized energy values. The energy of an electron in the n^{th} orbit of a hydrogen atom is $= -2.18 \times 10^{-18} \text{J}/n^2$

5. **The lowest energy state is the most stable**

The electron prefers to be in the lowest available energy level (ground state, $n=1$). Higher energy levels ($n>1$) are called **excited states**, and electrons tend to move back to lower levels by emitting photons.

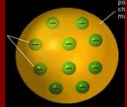
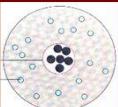
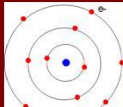
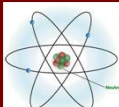
5. Quantum atomic model

- * It was put forward by Austrian physicist Erwin Schrödinger.
- * It stated that the behavior of atoms can be explained more clearly by treating them mathematically as matter waves.
- * This model is called quantum mechanical or wave mechanical model
- * *The Quantum Mechanical Model of the atom is based on the principles of quantum mechanics and describes the behavior of electrons in atoms.*

1.1.2. Key postulates of Quantum atomic model

1. Dual Nature of Electron: Electrons exhibit both particle and wave-like properties, as described by de Broglie's hypothesis.
2. Schrödinger Wave Equation: The behavior of electrons in an atom is governed by Schrödinger's wave equation, which describes electrons as wave functions (ψ) rather than fixed orbits.
3. Probability Distribution of Electrons: Electrons do not follow precise paths (as in the Bohr model); instead, their position is described by probability distributions known as atomic orbitals.
4. Quantization of Energy: Electrons can only occupy discrete energy levels (quantized states), as dictated by their wave functions.
5. Heisenberg's Uncertainty Principle: It is impossible to simultaneously determine both the position and momentum of an electron with absolute precision.
6. Orbitals and Quantum Numbers: The probability of finding electron, defined by 4 quantum numbers:
 - Principal Quantum Number (n) $\rightarrow$ Energy level
 - Azimuthal Quantum Number (l) $\rightarrow$ Orbital shape
 - Magnetic Quantum Number (m_l) $\rightarrow$ Orientation in space
 - Spin Quantum Number (m_s) $\rightarrow$ Electron spin ($\pm 1/2$)
7. Electron Density and Atomic Orbital: The probability of finding an electron in a particular region is given by $|\psi|^2$, which represents the electron density.

History of the Atomic Theory: Summary

					
1803	1897	1909	1913	1935	Today
solid Particle	electron	Proton, nucleus Random placed electrons	e- orbit nucleus: Planetary model, quanta	neutron	Quantum Atom theory
Dalton	Thomson	Rutherford	Bohr	Chadwick	Schrodinger and others

1.1.3. The difference between Bohr and Quantum Model

Bohr model Mo	Quantum model
Bohr model is an atomic model proposed by Niels Bohr (in 1915) to explain the structure of an atom.	Quantum model is an atomic model which is considered as the modern atomic model to explain the structure of an atom accurately.
The Bohr model explains the particle behavior of an electron.	Quantum model explains the wave-particle duality of an electron.
Bohr model can be applied for hydrogen atom but not for large atoms.	Quantum model can be used for any atom, including smaller ones and large, complex atoms.
Bohr model does not describe the exact shapes of every orbital.	Quantum model describes the all possible shapes an orbital can have.
Bohr model does not explain the Zeeman Effect (effect of magnetic field) or Stark effect (effect of electric field).	Quantum model explains the Zeeman and Stark effects precisely.
Bohr model does not describe the quantum numbers other than the principle quantum number.	Quantum model describes all four quantum numbers and the characteristics of an electron.

1.1.4. Limitation of the 5 Atomic Models

Atomic Model	Limitations
Dalton's Atomic Model	- Did not account for subatomic particles (electrons, protons, neutrons). - Could not explain isotopes (atoms of the same element with different masses). - Assumed atoms were indivisible and indestructible, which later disprove.
Thomson's Model	- Failed to explain the existence of a nucleus. - Could not account for the discrete spectral lines of elements. - Did not explain how electrons are arranged within the atom.
Rutherford's Atomic Model	- Did not explain why electrons do not spiral into the nucleus due to electrostatic attraction. - Could not account for atomic spectra and stability of atoms. - Did not describe electron energy levels.
Bohr's Model	- Could not explain spectra of multi-electron atoms. - Did not account for the wave nature of electrons.
Quantum Atomic Model	- Does not provide a simple, visual representation of atomic structure. - Based on complex mathematical probability functions (wave equations). - Difficult to conceptualize electron behavior compared to earlier models.

1.2. Questions with explanation

1. Which of a photon and an electron of de Broglie wave length having both with 1eV of energy
- A. The wave length of photon is greater than that of electron
 - B. The wave length of electron is greater than that of photon
 - C. The wave length of photon and electron are equal.
 - D. The information is not sufficient for comparing photon electron.

Brief:

- ❖ This question is describing the comparison of wave nature and particle nature of electron and photon.

Given $E=1\text{eV}$, and using $h=6.626\times 10^{-34}\text{Js}$, $c=3.0\times 10^8\text{ m/s}$, and $1\text{eV}=1.6\times 10^{-19}\text{J}$:

The energy of a photon is given by:

$$\lambda = \frac{hc}{E}$$

Rearranging for wavelength:

$$\lambda = \frac{hc}{E}$$

$$\lambda_{\text{photon}} = \frac{(6.626 \times 10^{-34})(3.0 \times 10^8)}{1.6 \times 10^{-19}}$$
$$\lambda_{\text{photon}} \approx 1.24 \times 10^{-6} \text{ m (or 1240 nm)}$$

Electron's Wavelength:

- ❖ or an electron, the de Broglie wavelength is: $\lambda = \frac{h}{p}$

where p is the momentum. The energy of a non-relativistic electron is: $E=p^2/2m$ and rearranging

$$p = \sqrt{2mE}$$

$$\lambda_{\text{electron}} = \frac{6.626 \times 10^{-34}}{5.4 \times 10^{-25}}$$

Using $m_e = 9.11 \times 10^{-31} \text{ kg}$ and $E = 1.6 \times 10^{-19} \text{ J}$:

$$p = \sqrt{2(9.11 \times 10^{-31})(1.6 \times 10^{-19})}$$

$$\lambda_{\text{electron}} \approx 1.23 \times 10^{-9} \text{ m (or 1.23 nm)}$$

$$p \approx 5.4 \times 10^{-25} \text{ kg m/s}$$

Therefore, Comparison of wave length of electron and photon is given as

$$\lambda_{\text{photon}} \approx 1240 \text{ nm}$$

$$\lambda_{\text{electron}} \approx 1.23 \text{ nm}$$

- $\lambda_{\text{photon}} > \lambda_{\text{electron}}$ implies that
- ✓ The wavelength of the photon is greater than that of the electron
 - ✓ Photon is more wave natured than electron.
 - ✓ Electron is more particle natured than photon.

2. What is the wave number of an electron traveling at the speed of $1.0 \times 10^7 \text{ m/s}$. (
 $m_e = 9.11 \times 10^{-31} \text{ Kg}$ & $h = 6.63 \times 10^{-34} \text{ J.S}$)
 A. 7.3×10^{-11} B. 7.3×10^{11} C. 1.3×10^{-12} D. 1.3×10^{-10}

Brief:

According to de Broglie's equation, $\lambda = h/mv$

Where

λ = de-Broglie wavelength and

$h = 6.626 \times 10^{-34} \text{ Js}$ = Planck's constant

m = mass of electron = $9.1 \times 10^{-31} \text{ kg}$

v = velocity = $1 \times 10^7 \text{ m/s}$

$\lambda = (6.626 \times 10^{-34} \text{ Js}) / (9.1 \times 10^{-31} \text{ Kg})(1 \times 10^7 \text{ m/s})$

Wave number = $1/\lambda$

$= 0.727 \times 10^{-10} \text{ m} = 7.3 \times 10^{-11} \text{ m}$

Wave number is the reciprocal of wave length.

🌈 Wave number = $1/7.3 \times 10^{-11} \text{ m}$

3. Arrange the various regions of the electromagnetic spectrum in the correct order in term of wave lengths

- ❖ Infrared > UV > visible > microwave > radio frequency
- ❖ Microwave > Infrared > visible > UV > radio frequency
- ❖ Radio frequency > Micro wave > Infrared > Visible > UV
- ❖ Visible > Infrared > UV > Microwave > radio frequency

Brief:

- ❖ Electromagnetic rays (EM rays) are certain type of energies in the form of waves.
- ❖ These waves are present all around us in our daily lives from the mobile radiations to scannings that doctors ask us to take.
- ❖ There are different variables that are used to describe EMR
- ✓ Wave length: Gamma rays < X Rays < UV Rays < Infrared Rays < Microwaves < Radiowaves
- ✓ Speed: Gamma rays > X Rays > UV Rays > Infrared Rays > Microwaves > Radiowaves
- ✓ Energy: Gamma rays > X Rays > UV Rays > Infrared Rays > Microwaves > Radiowaves
- ✓ Frequency: Gamma rays > X Rays > UV Rays > Infrared Rays > Microwaves > Radiowaves
- ✓ Wave number: Gamma rays > X Rays > UV Rays > Infrared Rays > Microwaves > Radiowaves

4. Which of the quantum numbers (n ,l, ml, ms) is permitted by the rule of quantum mechanics ?

- A. n =4 l=3 ml = -2 ms = $\pm \frac{1}{2}$
 B. n = 3 l =3 ml = -3 ms = $\pm \frac{1}{2}$
 C. n =3 l =2 ml = -4 ms = $\pm \frac{1}{2}$
 D. n = 1 l = 3 ml = -1 ms = $\pm \frac{1}{2}$

Brief:

Quantum No	Symbol	Description	Permissible values	Remarking
Principal quantum No	n	Energy and size of orbital	n=1,2, 3,4...	n is not to be zero
Angular Quantum No	l	Shape of orbital	0 to n-1	L=0 (s- orbital) L=1 (P-orbital) L=2 (d-orbital) L=3 (f-orbital)
Magnetic Quantum No	ml	Orientation of orbital	-L to +L	No of mL= 2L+1
Spin Quantum No	ms	Direction of electron	+/- 1/2	Only for electron

5. An arbitrary element X has two naturally occurring isotopes, information about the two isotopes is shown in the table below.

Isotopes	Atomic mass (amu)	Percent natural abundance (%)
X- 40	40	80
X-44	44	20

What is the average atomic mass of the element X?

- A. 41.5 B. 40.8 C. 42 D. 40

Brief:  Average atomic mass = $\frac{M_1f_1 + M_2f_2 + \dots + M_nf_n}{100}$

$$\text{Average atomic mass} = \frac{40\text{amu} (80) + (44 \text{amu}) (20)}{100}$$

$$\begin{aligned} \text{Average atomic mass} &= (40 \times 0.80) + (44 \times 0.20) \\ &= 32 + 8.8 = 40.8 \text{ amu} \qquad = 40.8 \text{ amu} \end{aligned}$$

6. Which of the following statements is accepted both in Dalton's and modern atomic theory?

- A. Atoms cannot be subdivided, created or destroyed.
- B. Atoms of the same elements may have different masses
- C. Atoms of a given elements are identical in mass and all other properties
- D. Atoms of different elements combine in simple whole number ratios.

Brief: Postulates of Dalton's Atomic Theory

- ✓ All matter is made up of tiny, indivisible particles called atoms.
- ✓ All atoms of a specific element are identical in mass, size, and other properties. However, atoms of different element exhibit different properties and vary in mass and size.
- ✓ Atoms can neither be created nor destroyed. Furthermore, atoms cannot be divided into smaller particles.
- ✓ Atoms of different elements can combine with each other in fixed whole-number ratios in order to form compounds.
- ✓ Atoms can be rearranged, combined, or separated in chemical reactions.
- ✓ Limitations of Dalton's Atomic TheoryIt does not account for subatomic particles:
Dalton's atomic theory stated that atoms were indivisible. However, the discovery of subatomic particles (such as protons, electrons, and neutrons) disproved this postulate.
- ✓ It does not account for isotopes: As per Dalton's atomic theory, all atoms of an element have identical masses and densities. However, different isotopes of elements have different atomic masses (Example: hydrogen, deuterium, and tritium).
- ✓ It does not account for isobars: This theory states that the masses of the atoms of two different elements must differ. However, it is possible for two different elements to share the same mass number. Such atoms are called isobars (Example: ^{40}Ar and ^{40}Ca).
- ✓ Elements need not combine in simple, whole-number ratios to form compounds:
Certain complex organic compounds do not feature simple ratios of constituent atoms.
Example: sugar/sucrose ($\text{C}_{11}\text{H}_{22}\text{O}_{11}$).
- ✓ The theory does not account for allotropes: The differences in the properties of diamond and graphite, both of which contain only carbon, cannot be explained by Dalton's atomic theory.

❖ Merits of Dalton's Atomic Theory

1. The law of multiple proportions, the law of conservation of mass, and the law of constant proportions are not violated by Dalton's atomic theory.
2. The theory provides a basis to differentiate between elements and compounds.

❖ Shortcomings of Dalton's atomic theory

- ✓ The theory did not account for the existence of subatomic particles (it suggested that atoms are indivisible).
 - ✓ suggesting that all atoms of an element must have identical masses and other properties
7. Light with wavelength of 400nm strikes the surface of metallic cesium (Cs) in a photo cell and the maximum K.E of the electron is 1.54×10^{-19} J. Calculate the longest wavelength of light

that is capable of ejecting of electron from the metal.

A. 3.34×10^{-19} m

B. 579m

C. 5.79nm

D. 579nm

Brief:

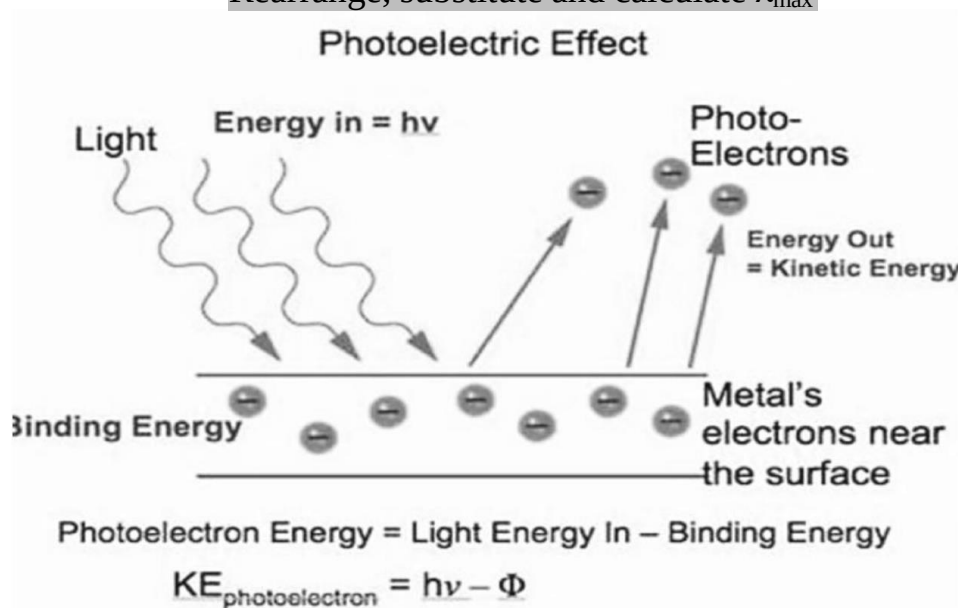
Given the information: $KE = 1.54 \times 10^{-19}$ J

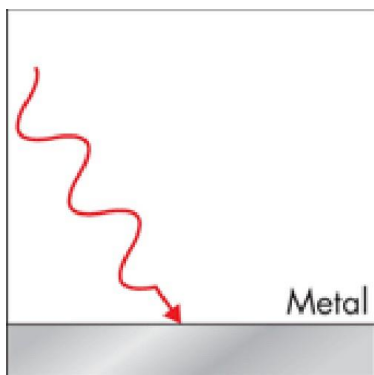
Wave length (λ) = 400nm = 4×10^{-7} m

Mass of electron = 9.11×10^{-34} kg

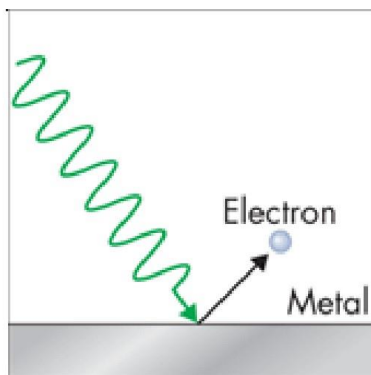
$$* h(c/\lambda) = \frac{1}{2}mV^2 + h(c/\lambda_{\max})$$

* Rearrange, substitute and calculate $\lambda_{\max}$

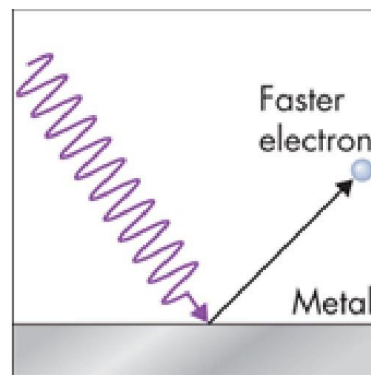




No electrons are ejected because the frequency of the light is below the threshold frequency.



If the light is at or above the threshold frequency, electrons are ejected.



If the frequency is increased, the ejected electrons will travel faster.

$$h\nu = KE + E_b$$

$$h\nu = \frac{1}{2} m_e v^2 + h\nu_0$$

8. Isotopes of the same element can differ in

- A. Mass B. Proton C. Atomic number D. Chemical property

Brief:

- Isotopes are atoms of the same element having the same
 - ✓ Number of proton
 - ✓ Number of electron
 - ✓ Atomic Number
 - ✓ Chemical properties
- But different
 - ✓ Mass number
 - ✓ Neutron number
 - ✓ Physical properties
- The chemical properties of atoms are determined from their electron number, so that Isotopes are similar in their electron number also similar in their chemical properties.
- Different isotopes exhibit nearly identical chemical behaviors.
- However, when it comes to physical properties of isotopes like the mass, melting or boiling point, density, and freezing point they all are different.
- The physical properties of any isotope mostly depend on the mass of each isotope.

9. Which of these sets orbitals is arranged in order of increasing energy?

- A. $3d < 4s < 4p < 5s < 4d$ C. $3d < 4s < 4p < 4d < 5s$
 B. $4s < 3d < 4p < 5s < 4d$ D. $4s < 3d < 4p < 4d < 5s$

Brief:

- As the principle quantum number (n) increases, the energy of the orbital also increases.
- Electrons are assigned to orbitals in order of increasing value of n+L.
- If two orbitals have the same n+L value, then the orbital with lower n value will have the lower energy. Hence, the increasing order of energy of the given orbitals will be:

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p...$$

10. Given the following orbitals: 3px, 2s, 4dxy, 3s, 4pz, 3py, 4s

Arrange the orbitals in the increasing order of their energy

- A. $2s < 3s = 3px = 3py < 4s = 4pz = 4dxy$
- B. $2s < 3s < 3px = 3py < 4s = 4pz = 4dxy$
- C. $2s < 3s < 3px = 3py < 4s < 4pz = 4dxy$
- D. $2s < 3s < 3px = 3py < 4s < 4pz < 4dxy$

Brief:

Orbitals are arranged with increasing energy on the basis of (n+l) rule. The higher the value of (n+l) for an orbital, the higher is its energy. Similarly, the lower the value of (n+l) for an orbital the lower is its energy. If two orbitals possess same (n+l) value, the orbital with lower (n+l) value will have the lower energy.

Let us find the energies of different orbitals.

➤ For 2s orbital:

- $n=2, l=0$
- $(n+l)=(2+0)=2$

➤ For 3s orbital:

- $n=3, l=0$
- $(n+l)=(3+0)=3$

➤ For 3px orbital:

- $n=3,$
 $l=1$

➤ 3py, 3px and 4s orbitals have same (n+l) value, but the n value of 3p orbital is lower than 4s orbital hence, it will have lower energy and also for the case of 3dxy and 4pz

➤ For 4dxy orbital:

- $n=4, l=2$
- $(n+l) = (4+2)=6.$

➤ For 3dxy orbital:

$$(n+l)=(3+1)=4$$

➤ For 3py orbital:

- $n=3, l=1$
- $(n+l)=(3+1)=4$

➤ For 4s orbital:

- $n=4, l=0$
- $(n+l)=(4+0)=4$

- $n=3, l=2$

➤ The order of increasing energies of orbitals is

$2s < 3s < 3p_x = 3p_y < 4s < 3d_{xy} < 4p_z < 4d_{xy}$

$$(n+l) = 3+2=5$$

➤ For $4p_z$ orbital:

- $n=4, l=1$
 $n+l) = (4+1)=5$

Some elements (like chromium and copper) show exceptions due to additional stability from half-filled or fully filled d-subshells. For example:

- **Chromium (Cr, Z = 24):** Expected: $[Ar] 4s^2 3d^4$, Actual: $[Ar] 4s^1 3d^5$
- **Copper (Cu, Z = 29):** Expected: $[Ar] 4s^2 3d^9$, Actual: $[Ar] 4s^1 3d^{10}$

There are three main principles of electron configuration

➤ **Hund's rule:**

1. According to Hund's Rule, if there are two or more degenerate orbitals (orbitals with the same energy) available, one electron must enter each one until they are all half full before pairing up.
2. Because every partially filled electron in an orbital of a subshell has the same spin, the orbital possess maximum stability
3. Electrons will fill all the degenerate orbitals (equal in energy) with parallel spins (both arrows up or down) first before pairing up in one orbital. We can also formulate it as the lowest energy configuration for an atom is the one having the maximum number of unpaired electrons within the same energy sublevel.

4. Hund's rule is applicable for configuration with degenerating orbitals

Degenerating orbitals : Those are orbitals with the same energy

s.....No degenerating orbital (only s)

p.....3 degenerating orbitals (p_x, p_y, p_z)

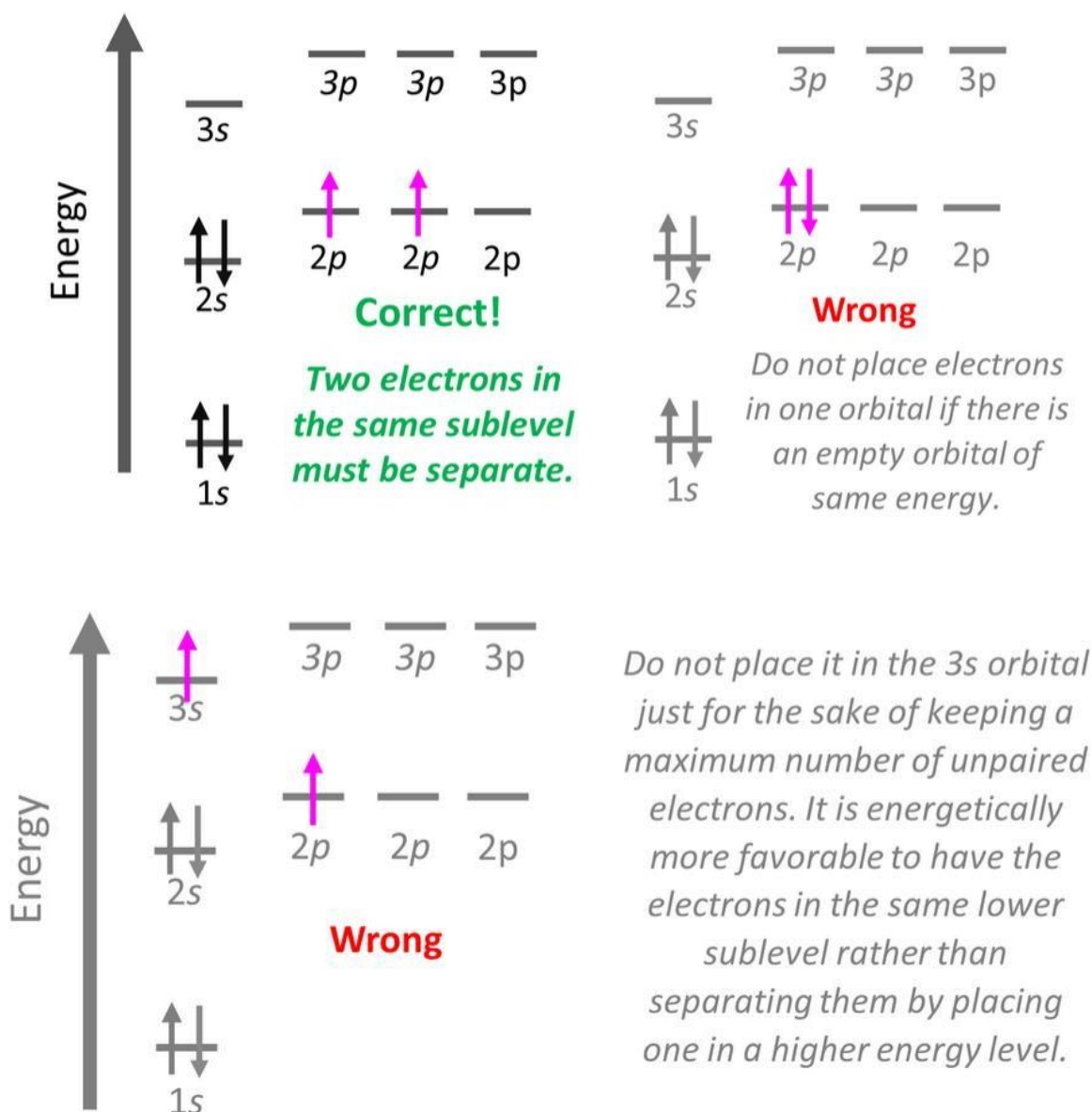
d.....5 degenerating orbitals ($d_{xy}, d_{xz}, d_{yz}, d_{x^2-y^2}, d_{z^2}$)

➤ **Pauli Exclusion Principle:**

1. According to the Pauli Exclusion Principle, no two electrons in a single atom will have the same set of quantum numbers (n , l , m_l , and m_s). Simply said, each electron should have or exist in its own unique state (singlet state).
2. The two electrons in the same orbital must have opposing spins or be in an antiparallel configuration.
3. According to the Pauli Exclusion Principle which states that no two electrons in an atom can have the same *four* quantum numbers. So, if the electrons are in the same orbital, they must have the same n , l , m_l values, and therefore, the only one that can be different is the m_s which is the spin of the electron shown by the position of the arrow.

➤ **Aufbau principle:**

1. According to this, electrons are placed into atomic orbitals in ascending sequence of orbital energy. The Aufbau principle states that atomic orbitals that are available and have the lowest energy levels get filled before those that have higher energy levels.
2. The sequence of energy levels is: $1s$, $2s$, $2p$, $3s$, $3p$, $4s$, $3d$, $4p$, $5s$, $4d$, $5p$, $6s$, $5d$, $6p$, $7s$, $6d$, $7p$, $8s$, $7f$, $8d$, $9p$, $10s$, $9f$, $10d$, $11p$, $12s$, $11d$, $12f$, $13p$, $14s$, $13d$, $14f$, $15p$, $16s$, $15d$, $16f$, $17p$, $18s$, $17d$, $18f$, $19p$, $20s$, $19d$, $20f$, $21p$, $22s$, $21d$, $22f$, $23p$, $24s$, $23d$, $24f$, $25p$, $26s$, $25d$, $26f$, $27p$, $28s$, $27d$, $28f$, $29p$, $30s$, $29d$, $30f$, $31p$, $32s$, $31d$, $32f$, $33p$, $34s$, $33d$, $34f$, $35p$, $36s$, $35d$, $36f$, $37p$, $38s$, $37d$, $38f$, $39p$, $40s$, $39d$, $40f$, $41p$, $42s$, $41d$, $42f$, $43p$, $44s$, $43d$, $44f$, $45p$, $46s$, $45d$, $46f$, $47p$, $48s$, $47d$, $48f$, $49p$, $50s$, $49d$, $50f$, $51p$, $52s$, $51d$, $52f$, $53p$, $54s$, $53d$, $54f$, $55p$, $56s$, $55d$, $56f$, $57p$, $58s$, $57d$, $58f$, $59p$, $60s$, $59d$, $60f$, $61p$, $62s$, $61d$, $62f$, $63p$, $64s$, $63d$, $64f$, $65p$, $66s$, $65d$, $66f$, $67p$, $68s$, $67d$, $68f$, $69p$, 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11. Among the following, which color corresponds to light of the highest Frequency?

- A. Green B. Yellow C. Red D. Blue

Brief:

- ✚ The increasing order of colors depends on different parameters
 - frequency
 - Energy
 - Wave number
 - Wave length :
- red < orange < yellow < green < blue < indigo < violet
- red > orange > yellow > green > blue > indigo > violet

12. Which of these sets of n, l, ml, ms quantum numbers can be used to describe an electron in a 2p_x orbital. A. 2, 1, -1, +1/2 B. 2, -2, 0, 1/2 C. 2, 2, 1, 1/2 D. 2, 2, 1, -1/2

Brief:

✚ The four permissible quantum numbers n, l, ml, ms for 2p orbital are: 2, 1, -1, +1/2

13. The wave length of an electron moving at 1/1000 the speed of light is

A. 2.43pm B. 2.43×10^{-9} cm C. 2.43m D. 2.43nm

Brief:

✚ Given the speed of electron = $1/1000 (3 \times 10^8) = 3 \times 10^5$ m/s

Wave length = $h/mv = (6.6263 \times 10^{-34} \text{ Js}) / (9.11 \times 10^{-31} \text{ Kg})(3 \times 10^5 \text{ m/s})$

14. The de Broglie wavelength of an electron with a velocity 10.% of the speed of light and tennis ball with 55 g served at 35 m/s are _____ and _____ respectively.

A. 2.42×10^{-11} m and 3.44×10^{-34} m C. 2.42×10^{-14} m and 3.44×10^{-30} m
B. 2.42×10^{-11} m and 3.44×10^{-18} m D. 2.42×10^{-11} m and 3.44×10^{-25} m

Brief:

- ❖ To calculate the **de Broglie wavelength** of electron and a tennis ball for that matter, just use the Equation

$$p = \frac{h}{\lambda} \Rightarrow m \cdot v = \frac{h}{\lambda} \Rightarrow \lambda = \frac{h}{m \cdot v}$$

$$\lambda_{\text{electron}} = \frac{6.626 \cdot 10^{-34} \cancel{\text{m}^2 \text{ kg s}^{-1}}}{9.1094 \cdot 10^{-31} \cancel{\text{kg}} \cdot 3 \cdot 10^7 \cancel{\text{m s}^{-1}}}$$

$$\lambda_{\text{electron}} = 2.42 \cdot 10^{-11} \text{m}$$

→ The more wave

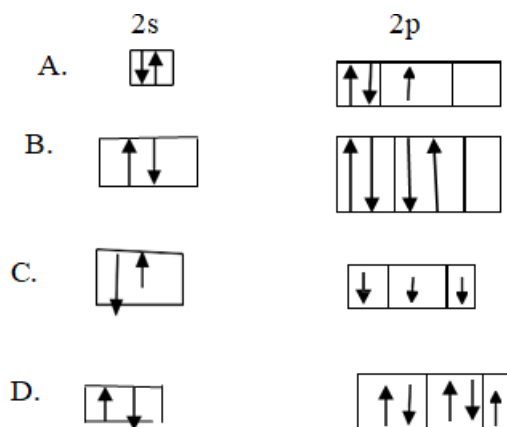
Now for the tennis ball

$$\lambda_{\text{tennis}} = \frac{6.626 \cdot 10^{-34} \cancel{\text{m}^2 \text{ kg s}^{-1}}}{55 \cdot 10^{-3} \cancel{\text{kg}} \cdot 35 \cancel{\text{m s}^{-1}}}$$

$$\lambda_{\text{tennis}} = 3.44 \cdot 10^{-34} \text{m}$$

→ The more particle

15. The orbital diagram in which both Pauli's exclusion principle and Hund's rules are violated?



Brief:

- ❖ In degenerating orbitals electrons are distributed singly before enter to double
- ❖ The two electrons in the same orbit should be in opposite spin
- ❖ Electrons those are in degenerating orbitals should be in the same spin direction

16. An element with the general electron configuration for its outermost electron of ns^2np^6 would be in which group. A. IIA B. IIIA C. VIIA D. VIA E. VIIIA

Brief:

- ❖ An electronic configuration, the arrangement of electrons in orbitals around an atomic nucleus. The electronic configuration of an atom in the quantum-mechanical model is stated by listing the occupied orbitals, in order of filling, with the number of electrons in each orbital indicated by superscript.
- ❖ The valance shell electron **configuration of representative elements is: ns^2np^6**

Group	Outermost shell electron configuration	Block
IA	ns^1	s-block
II	ns^2	s-block
III	ns^2np^1	p-block
IV	ns^2np^2	p-block
V	ns^2np^3	p-block
VI	ns^2np^4	p-block
VII	ns^2np^5	p-block
VIII	ns^2np^6	p-block

- ❖ The general electronic configuration of transition element is **$(n-1)d^{1-10}ns^{1-2}$** .

17. An electron has spin quantum number, $m_s = +1/2$ and magnetic quantum number, $m_l = +1$ cannot be presented

- A. p-orbital B. d-orbital C f-orbital D. s- orbital

Brief:

❖ From The given values:

- $m_s = +1/2$, it tells about the direction of electron where any of orbitals
- $m = +1$, it tells the orientation of orbital that can be valued on p,d,f orbitals
- There is only one m_l value of s-orbital, ($m_l=0$) not $m_l=1$.

18. Calculate the energy for energy level 3 in hydrogen

- A. -2.18×10^{-18} J B. -2.42×10^{-19} J C. -2.42×10^{-18} J D. -5.45×10^{-19} J

Brief:

$$E = -2.18 \times 10^{-18} \text{ J } (Z^2/n^2)$$

$$E = -2.18 \times 10^{-18} \text{ J } (1^2/3^2)$$

$$E = -2.18 \times 10^{-18} \text{ J } (1/9)$$

$$E = -2.42 \times 10^{-19} \text{ J}$$

19. Which one of the following statements is not true about quantum Mechanical models of the atom?

- A. In Bohr's Theory, electrons were thought to as moving around the nucleus in circular orbital. But The Bohr's theory could explain the atoms with more than one electron.
- B. In quantum mechanics the electron moves around the nucleus but the exact path they take can't be determined
- C. The Heisenberg's uncertainty principle states that it is impossible to determine accurately both the momentum and position of a particle like an electron simultaneously.
- D. Erwin Schrödinger developed complex mathematical equation that relates the wave properties associated with electrons to their energies.

Brief:

❖ **Louis de Broglie** proposed that all particles could be treated as matter waves with a wavelength

$$\lambda = \frac{h}{mv}$$

❖ **Erwin Schrödinger** proposed the quantum mechanical model of the atom, which treats electrons as matter waves.

- ❖ Schrödinger's equation, $\hat{H}\psi = E\psi$ can be solved to yield a series of wave function, each of which is associated with an electron binding energy. The square of the wave function, represents the *probability* of finding an electron in a given region within the atom.
 - ❖ Electrons have an intrinsic property called spin, and an electron can have one of two possible spin values: (spin-up or spin-down).
 - ❖ Any two electrons occupying the same orbital must have opposite spins.
 - ❖ The Heisenberg's uncertainty principle states that it is impossible to determine accurately both the momentum and position of a particle like an electron simultaneously.
 - ❖ In Bohr's Theory, electrons were thought to be moving around the nucleus in circular orbital. But Bohr's theory could not explain the atoms with more than one electron.
20. The wavelength of light of wavelength $4000\text{\AA}$ falls on the surface of cesium. Calculate the energy of the photoelectron emitted. The critical wavelength for photoelectric effect in cesium is $6600\text{\AA}$.
- A. 1.96×10^{-19} B. 1.96×10^{-19} C. 1.96×10^{-19} D. 4.76×10^{-18}

Brief:

$$\begin{aligned}
 KE &= h\nu - h\nu_0 \\
 &= \frac{hc}{\lambda} - \frac{hc}{\lambda_0} \\
 &= hc \left(\frac{1}{4000} - \frac{1}{6600} \right) \\
 &= 6.626 \times 10^{-34} \times \frac{3 \times 10^8}{10^{-10}} \times \frac{2600}{4000 \times 6600} \\
 &= 1.9577 \times 10^{-19} \text{ J}
 \end{aligned}$$

Energy of the photoelectrons emitted = $1.9577 \times 10^{-19} \text{ J}$

21. Quantum mechanics suggests that the total energy of an electron in an atom is quantized. This is the work of: A. Schrodinger B. Niels Bohr C. Heisenberg D. De Broglie.

Brief:

I. **Schrodinger**... Describing the energy and position of the electron in the function of space and time by suggesting the total energy of an electron in an atom is quantized. He developed the equation that gives the probability of finding an electron.

II. **Niels Bohr**.... Electrons revolve around the nucleus in stable orbits without emission of radiant energy. But when electrons jump from one energy level to

another it either emit (from higher to lower) or absorb (from lower to higher) energy

III. **Heisenberg**.... Impossible to identify the exact position and Momentum of electron simultaneously

IV. **De Broglie**.... predicted that all matter can behave as waves as well as particles (dual nature of particle)

22. Which of the following has the smallest radius?

- A. Ne B. O^{2-} C. Mg^{+2} D. F^- E. Na^+

Brief:

- ❖ Atomic size is the distance between the centre of the nucleus of an atom and its outermost shell.
- ❖ Across a period, Atomic size decrease due to the more outer most shell electron
- ❖ Down a group, Atomic size increase due to the more number of shell available.
- ❖ Atomic size is also affected by the magnitude of the charge; as the magnitude of positive charge increase size of ion decrease.
- ❖ Anions are larger than the corresponding neutral atom, since adding electrons increases the number of electron-electron repulsion interactions that take place.
- ❖ Cations are smaller than the corresponding neutral atoms, since the valence electrons, which are furthest away from the nucleus, are lost.
- ❖ Size: anion > atom > cation: $O^{2-} > F^- > Ne > Na^+ > Mg^{+2}$

Practice: Arrange the following species in terms of Atomic size, ionization energy, electron negativity and electron affinity.



23. Which of the following statement is true about quantum number?

- A. The number of electron in an atom can have when $n=4$, and $m_s = \pm 1/2 = 32$
- B. The number of electron in an atom can have when $n, 4$ and $l = 4 = 18$
- C. The number of electron in an atom can have when $n = 6, l = 3$ and $m_l = -2 = 14$
- D. The number of electron in an atom can have when $n = 5, = 25$

Brief:

- ❖ At $n=4$: The number of orbitals n^2 (4s, 4p, 4d, 4f) . The number of electron also $=2n^2$ so The number of electron in an atom can have when $n= 4$ is 32
- ❖ At $n=4$, $l = 0$ to $n-1$ (not 4).
- ❖ At $n=6$ and $l=3$ $ml= -2$, Only two electrons should occupy
- ❖ The number of electron in an atom can have when $n = 5$, = 25

24. In what specific sub-shell will an electron having the quantum number $n = 3$, $l = 2$ and $ml = -1$ be found.

- A. 3f B. 3s C. 3p D. 3d

Brief:

$n=3$ at 3rd energy level

$l=2$ d – orbital

$ml=2$ orientation

from the information given: subshell is =3d

25. Which of the following electron transition required the smallest energy to be absorbed by the hydrogen atom?

- A. From $n = 3$ to $n=4$ C. from $n= 4$ to $n=5$
 B. From $n= 6$ to $n = 7$ D. from $n= 7$ to $n= 8$

Brief

- ❖ Generally, The electronic transition in a hydrogen atom associated with the smallest change in energy is $n = 2$ to $n = 1$. When an electron moves from a higher energy level ($n = 2$) to a lower energy level ($n = 1$), it releases the least amount of energy.
- ❖ The energy of radiation after a transition is given by

$$\Delta E = -2.18 \times 10^{-18} \text{ J} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

26. Using the expression for angular momentum calculate the velocity of the electron in $n=3$ where $r=n^2a_0$.

- A. $7.28 \times 10^5 \text{ m/s}$ B. $4.77 \times 10^{-10} \text{ m/s}$
 B. C. $4.77 \times 10^{-10} \text{ m}$ D None

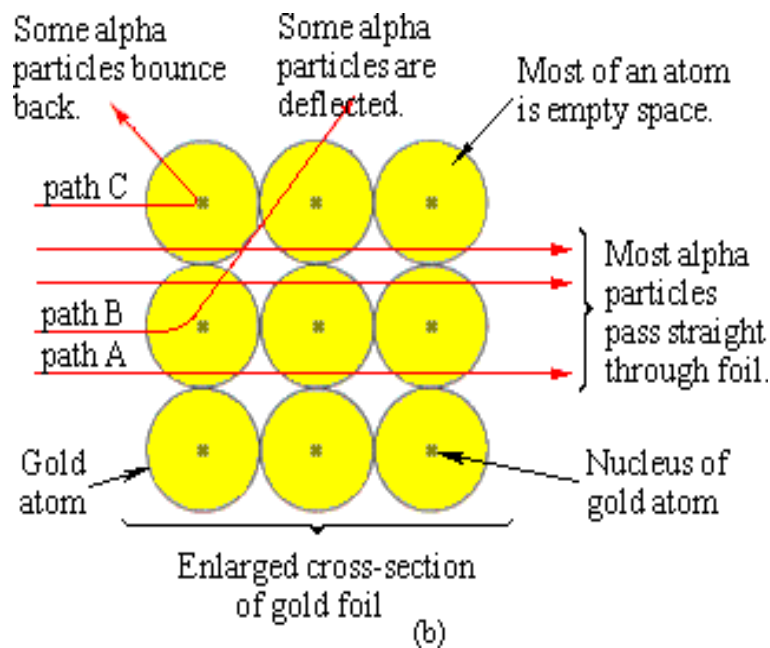
Brief:

- ❖ The angular momentum of an electron by Bohr is given by $mvr = nh/2\pi$.
 The value of r calculated by: $r = n^2a_0$
 $v = nh/2\pi mr = \frac{nh}{2\pi m(n^2a_0)}$
 $(n=3, h= 6.626 \times 10^{-34} \text{ m}^2 \text{ kg / s}, m_e = 9.11 \times 10^{-31} \text{ Kg}, a_0=0.0529\text{nm})$

27. Which conclusion are bases on the Rutherford's gold foil experiments and the resulting models of the atom?

- A. An atom is mainly empty space and the nucleus has a positive charge
- B. An atom is mainly empty space and the nucleus has a negative charge
- C. An atom has hardly any empty space, and the nucleus has a positive charge.
- D. An atom has hardly any empty space, and the nucleus has a negative charge.

- Brief:** ❖ The gold-foil experiment showed that the atom consists of a small, massive, positively charged nucleus with the negatively charged electrons being at a great distance from the center.
- ❖ The gold foil experiment led to the conclusion that each atom in the foil was composed mostly of empty space because most alpha particles directed at the foil
 - ❖ An atom is mostly empty space with a dense, positively charged nucleus.
 - ❖ An atom is composed of at least three types of subatomic particles.



28. The threshold frequency for metallic potassium is $6 \times 10^{14} \text{ Hz}$ calculate the velocity of the ejection electron will have a wave length of light shining on the potassium surface is 360 nm.

- A. $8.33 \times 10^{14} \text{ m/s}$ B. $6.73 \times 10^7 \text{ m/s}$ C. $0.529 \times 10^{-6} \text{ m/s}$ D. $5.29 \times 10^5 \text{ m/s}$

Brief:

➤ Given that

- $h = 6.63 \times 10^{-34} \text{ Js}$
- $V_0 = 6 \times 10^{14} \text{ Hz}$
- $\lambda = 360 \text{ nm} = 3.6 \times 10^{-7} \text{ m}$
- $3 \times 10^8 \text{ m/s}$

➤ Required : Velocity of an electron ejecting from surface of metal

- $E = hV_0 + \text{KE}$
- $h\nu = hV_0 + \text{KE}$
- $\text{KE} = h\nu - hV_0$
- $\frac{1}{2}mV^2 = h\nu - hC$
- $V^2 = \frac{2h(V - V_0)}{m}$
- $\sqrt{\left(\frac{2h(V - V_0)}{m}\right)}$
- $\sqrt{\left(\frac{2h\left(\frac{C}{\lambda} - V_0\right)}{m}\right)}$

Where:

$$V = C / \lambda$$

$$= (3 \times 10^8 \text{ m/s}) / (3.6 \times 10^{-7} \text{ m})$$

Substitute all values and find the last answer.

29. What is the velocity of a 300g ball having a De Broglie wave length of $0.06 \times 10^{-10} \text{ m}$?

- A. $6.63(0.06135)(0.300) \times 10^{-34}$ C. $0.06135 \times 10^{-10}(0.300) / (6.63 \times 10^{-34})$
 B. $6.63 \times 10^{-34} / 0.06135 \times 10^{-10}(0.300)$ D. $6.63 \times 10^{-34}(0.300) / 0.06135 \times 10^{-10}$

Brief:

- Given the de Broglie equation : $\lambda = h/mv$
- $V = h/m\lambda = (6.63 \times 10^{-34} \text{ Js}) / (0.06 \times 10^{-10} \text{ m})(0.300 \text{ Kg})$

30. Which of the following statements is NOT true of transition elements?

- They are all metals C. Most of their ions are colored
 A. Few of them form complex D. most of them contain unpaired electrons
 E. Most of them have incompletely filled orbitals.

Brief:

➤ Properties of transition elements

✚ It is useful, at the beginning, to identify the physical and chemical properties of transition elements which differ from main group elements (s-block).

✚ Common properties of transition elements include:

- ❖ have large charge/radius ratio;
- ❖ are hard and have high densities
- ❖ have high melting and boiling points
- ❖ form compounds which are often paramagnetic
- ❖ show variable oxidation states;
- ❖ form coloured ions and compounds;
- ❖ form compounds with profound catalytic activity;
- ❖ Form stable complexes

31. The halogens are best described by which of the following statements?

- A. They are members of group VIII in the periodic table
- B. They are all gases at room temperature and pressure
- C. They are all oxidizing agents
- D. Most of them are colorless

The halogens are all **non-metals**. They show many of the physical properties typical of non-metals.

Brief:

- ✓ They are poor conductors of heat and electricity.
- ✓ When solid, they are dull and brittle.
- ✓ They have low melting and boiling points.
- ✓ They are members of group VII in the periodic table
- ✓ Melting and boiling points increase as you go down the group due to the atoms get larger
- ✓ They have high electronegativity values.
- ✓ Chlorine, bromine, and iodine are all sparingly soluble in water.
- ✓ Halogens tend to react by gaining an electron. This means that they act as oxidising agents and are reduced themselves

32. When an electron jumps from secondary energy level to third energy levels which of the following phenomena are correct?

- A .Energy is release B. energy is absorbed C. The system is exothermic D. all

Brief:

➤ According to Bohr's model

- ✚ In an atom, the electrons revolve around the nucleus in a definite circular path known as the orbits.
- ✚ It is possible to determine the exact location of electron that far from the nucleus by: $r = a_0 n^2$,
- ✚ The energy of electron at each stationery state can be calculated as: $E = -(2.18 \times 10^{-18})/n^2$
- ✚ When electron revolving around the nucleus energy not radiant
- ✚ The energy levels are denoted by a positive integer ($n=1, 2, 3 \dots$) which is known as the quantum number.
- ✚ The electrons gaining the required amount of energy move from their lower energy level to a higher energy level and vice versa.

33. Which of the following statements are NOT correctly described about Bohr's model?

- A. Angular momentum of an electron in a hydrogen atom is quantized.
- B. In an atom only those orbits are permitted in which angular momentum of the electron is a natural number multiple of 2π
- C. Electron revolves around nucleus only in those orbits for which the angular momentum is an integral multiple of $h/2\pi$.
- D. The Coulomb force is responsible for keeping the electron in orbit around the positive nucleus

Brief:

➤ The following are important assumption

- ✚ Electrons are only stable in certain orbits
- ✚ Atoms only emit electromagnetic radiation when electrons transition to lower energy states
- ✚ Atoms only absorb electromagnetic radiation when electrons transition to higher energy states
- ✚ The Coulomb force is responsible for keeping the electron in orbit around the positive nucleus
- ✚ Angular momentum of an electron in a hydrogen atom is quantized.
- ✚ In an atom only those orbits are permitted in which angular momentum of the electron is a natural number multiple of $h/2\pi$, $mvr = h/2\pi$
- ✚ The Coulomb force is responsible for keeping the electron in orbit around the positive nucleus

34. Which of the following equations is not the assumption made by the Bohr model

A. $mvr = \frac{nh}{2\pi}$

$r = n^2 (0.529) \text{ \AA}$

C. $E = mc^2$

D. $E = -2.18 \times 10^{-18} \text{ J } (1/n^2)$

Brief:

A. $mvr = \frac{nh}{2\pi}$	$r = n^2 (0.529) \text{ \AA}$	C. $E = mc^2$	D. $E = -2.18 \times 10^{-18} \text{ J}$
Angular momentum of electron	Radius of electron from the nucleus	Albert Einstein equation about energy	Energy of electron on each stationary state

35. Which one of the following electromagnetic-radiations has the highest energy?

A. X-rays

B. Uv-ray

C. Gamma ray

D. Microwave

Brief:

✚ The order of EMR based on different parameters

❖ **Based on frequency/energy/wave number** is:-

Gamma rays > X-rays > ultra-violet radiation > infrared radiation

❖ **Based on Wave length:**

Gamma rays < X-rays < ultra-violet radiation < infrared radiation

36. In general, as you go across a period in the periodic table from left to right:

(1) The atomic radius

(2) The electron affinity

(3) The first ionization energy

A. Decreases, increasingly, decreases

C. Decreases, increasingly, increases

B. Increases, increasingly, decreases

D. Increases, decreasingly, increases

Brief

Periodic properties	Trend		Reason
	Across a period	Down a group	
Nuclear charge	increase	increase	Atomic number
Effective nuclear charge	increase	constant	Outer most shell electron
Atomic size	decrease	increase	Nucleus-valence electron interaction
Ionization energy	increase	decrease	Nucleus-valence electron interaction
Electron negativity	increase	decrease	Ability of attracting electron
Electron affinity	increase	decrease	Energy released on adding of electron
Metallic character	decrease	increase	Position of metals in periodic table
Non-metallic character	increase	decrease	Position of non-metals in periodic table

37. Electro-negativity is a concept that is useful along with other concepts in:

- A. Predicting the polarity of a bond
- B. Deciding how many electrons are involved in a bond
- C. Formulating a statement of the octet rule
- D. Determining the charge on a polyatomic ion

Brief

- ✓ The polarity of molecules is based on the electronegativity of molecules.
- ✓ The more electronegativity difference between atoms the more polar
(eg. HBr, HI, HCl, HF)
- ✓ The least electronegativity difference between atoms tends to have less polar (e.g CH₄).
- ✓ The first most electronegative atoms are termed as: FONClBrICH
(F > O > N > Cl > Br > I > C > H)

38. In photoelectric effect, the number of photoelectrons emitted per unit area per unit depends on

- A. Frequency of the incident radiation
- B. Wavelength of the incident radiation
- C. Intensive of the incident radiation
- D. Wave number of the incident radiation.

Brief:

In Photoelectric effect, Energy of emitted photoelectrons depends on the frequency of incident electromagnetic radiation and the number of photo electrons emitted per second depends on intensity of the incident radiation.

The number of electrons ejected is proportional to the intensity or brightness of light in experiment of photoelectric effect.

➤ Necessary Conditions for Photoelectric Effect

- * The threshold frequency has to be attained which is expressed as $E = h\nu_0$.
- * Meet or exceed the metal's work function.
- * Decrease wavelength.
- * Increase frequency.
- * Increasing intensity of light.

39. In an atom the maximum number of electrons having the quantum number $n=6$, $l=3$ and $m_s=\pm 1/2$

- A. 14 B. 7 C. 32 D. 6 E. 16

Brief:

The maximum number of electron at $n=6$ (6^{th} energy level) and $l=3$ (f-orbital) is 14.

The five main atomic models are:

Dalton's atomic model $\rightarrow$ Thomson's model $\rightarrow$ Thomson's model $\rightarrow$

Rutherford model $\rightarrow$ Bohr's Model $\rightarrow$ Quantum mechanical model

40. Which of the following is a correct interpretation of the results of Rutherford's experiments in which gold atoms were bombarded with alpha particles?

- A. Atoms have equal numbers of positive and negative charges.
- B. The positive charge of an atom is concentrated in a small region
- C. Neutrons and protons in atoms have nearly equal mass.
- D. Electrons in atoms are arranged in shells.

Brief:

- ✓ Rutherford's gold foil experiment showed that alpha particles were deflected by a small, dense region at the center of the atom, which he concluded must be positively charged (the nucleus).
- ✓ While neutrons and protons do have nearly equal mass, it is not the conclusion of Rutherford's experiment.

41. In Bohr's stationary orbits:

- A. Electrons do not move
- B. Electrons move emitting radiation
- C. Energy of the electron remains constant
- D. Angular momentum of the electron is $h/2\pi$

Brief:

According to Bohr Model:

- ✓ Electrons do move, but they do so in specific orbits without radiating energy.
- ✓ Electrons in stationary orbits do not emit radiation, and their energy is constant while they are in these orbits.
- ✓ Electrons only emit radiation when transitioning between orbits, not while stationary in a given orbit.
- ✓ Angular momentum of the electron is $h/2\pi$. An electron in an atom could revolve in certain stable orbits without the emission of radiant energy.

- ✓ When electron makes a transition from one of its stable orbit to lower orbit then a photon is emitted with energy $h\nu = E_f - E_i$
- ✓ Bohr's model is specifically applicable to hydrogen-like atoms (atoms with a single electron, like hydrogen). It does not apply to atoms with more than one electron.

42. Which series of hydrogen spectrum corresponds to ultraviolet region?

- A. Balmer series B. Brackett series C. Paschen series D. Lyman series

Brief:

- ✓ The Balmer series..... corresponds to visible light.
- ✓ The Paschen series..... corresponds to infrared light.
- ✓ The Brackett series..... corresponds to infrared light.
- ✓ The Lyman series..... corresponds to ultraviolet light and involves transitions to the $n=1$ level.

43. What is the key postulate of the Bohr model of the hydrogen atom?

- A. Electrons are in elliptical orbits around the nucleus.
 B. Electrons have definite orbits with quantized energy levels.
 C. Electrons are found uniformly in a sphere around the nucleus.
 D. Electrons do not follow any specific paths.

Brief:

- ✓ Bohr's original model has circular orbits, not elliptical.
- ✓ The key postulate of the Bohr model, where
- ✓ "Electrons have definite orbits with quantized energy levels" is the key postulate of the Bohr model, where electrons occupy discrete energy levels.
- ✓ "Electrons do not follow any specific paths" is more related to quantum mechanical models.

44. Choose the correct statement(s) that are proposed by Neils Bohr, regarding the model of an atom.

- A. Electrons absorb energy when they move to higher energy orbits.
 B. Electrons emit energy when they move to lower energy orbits.
 C. Electrons absorb energy when they move to lower energy orbits.
 D. Electrons emit energy when they move to higher energy orbits.

Brief:

- ✓ According to Bohr Model, electrons neither emit nor absorb energy when they are rotating around the nucleus regarding to fixed stationary state.
- ✓ When Electrons are jump from one energy level to another, it emit or absorb energy. i.e
 - From lower to higher energy level.....Absorb energy

- From higher to lower energy level Emit energy

45. Which of the following conclusions regarding the structure of atom is based on Rutherford's α -particle scattering experiment?

- A. The positive charge is concentrated in a very small volume of the atom
- B. The positive charge is scattered with the electrons throughout the atom
- C. The volume occupied by the nucleus is half of the volume of atom
- D. Most of the space in the atom is occupied by the neutron

Brief:

Rutherford concluded the following summary about an atom after his experiment

- ✓ Atom has a dense, positively charged nucleus: This was based on the observation that most α -particles passed through the gold foil with little deflection, but some were deflected at large angles, and a very small fraction even rebounded. This indicated that the mass and positive charge of the atom are concentrated in a tiny central core (the nucleus).
- ✓ The atom is mostly empty space: Since most of the α -particles passed straight through the foil, it was concluded that atoms are mostly empty space, with the nucleus taking up a very small volume in the center.
- ✓ Size of the nucleus: The nucleus was found to be incredibly small compared to the overall size of the atom. The deflections of the α -particles were caused by their interaction with the dense nucleus, but most of the atom's space was empty.
- ✓ The existence of protons: While Rutherford's result didn't confirmed directly to the existence of atom, it laid the groundwork for the future understanding that the nucleus is made up of positively charged protons

46. Among the following pairs of compounds, the one that illustrates the law of multiple proportions is

- | | |
|---|---|
| A. NH_3 and NCl_3 | C. CS_2 and FeSO_4 |
| B. H_2S and SO_2 | D. CuO and Cu_2O |

Brief:

- ✓ The Law of Multiple Proportions is formulated by John Dalton. It states that when two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in simple whole-number ratios.
- ✓ CuO and Cu_2O are both compounds of copper and oxygen, but with different stoichiometry of copper.

- ✓ CuO has 1 copper atom for each oxygen atom, while Cu₂O has 2 copper atoms for each oxygen atom. This is an example of the law of multiple proportions while Both compounds consist of copper and oxygen, but the ratio of copper to oxygen is a simple whole number ratio: 1:1 in CuO and 2:1 in Cu₂O

47. What occurs when an electron moves from high energy level to a low one?

- A. another electron goes from a low energy level to a higher
- B. The atom moves faster
- C. Colored light is given off
- D. This process is not possible

Brief:

- ✓ When an electron moves from a higher energy level to a lower one, the energy difference between these levels is released in the form of light. This light can be of a specific color, depending on the energy difference. This process is known as **emission**, and the color corresponds to the wavelength of the emitted light.

48. Which of the following is true according to the Planetary Model of the atom?

- A. Electrons follow fixed paths around the nucleus.
- B. The atom is indivisible and uniform throughout.
- C. Electrons are found in orbitals, which are regions of space around the nucleus.
- D. The nucleus is composed of electrons and protons.

Brief:

- ✓ According to the Planetary Model of the atom, proposed by Niels Bohr, electrons move in fixed, circular orbits around the nucleus, much like planets orbit the sun.

49. Which statement is incorrect?

- A. Orbital is a region in an atom where an electron can be found.
- B. An electron can absorb energy when it jumps to a higher energy level.
- C. An electron can emit energy when it jumps to a higher energy level.
- D. Filling of electrons in an atom starts from a low energy level to the highest energy level.

Brief:

An electron absorbs energy when it jumps to a higher energy level While, When an electron moves to a lower energy level, it emits energy in the form of light (photons).

50. Which of the following statement is correct?

- A. Ultraviolet radiation has a longer wavelength than infrared radiation
- B. Microwave radiation possesses more energy than infrared radiation
- C. Infrared radiation has a shorter wavelength than visible light
- D. Infrared radiation has a lower wavenumber than visible light

Brief:

- ✓ Ultraviolet radiation has a shorter wavelength than infrared radiation.
- ✓ Microwave radiation possesses less energy and low frequency than infrared radiation
- ✓ Infrared radiation has a longer wavelength than visible light
- ✓ The wavenumber is inversely proportional to the wavelength. Since infrared radiation has longer wavelengths than visible light, it also has a lower wavenumber.

1.3. Self-Practice Questions

51. Which of the following electromagnetic waves is primarily used in medical imaging to view bones inside the body?

- A. Microwaves B. X-rays C. Infrared rays D. Ultraviolet ray

52. Which property of electromagnetic radiation determines how much energy a wave can deliver

- A. Frequency B. Speed C. Wavelength D. Polarization

53. Which of the following correctly represents the four quantum numbers that define an electron in an atom?

- A. Principal Quantum Number (n), Angular Momentum Quantum Number (l), Magnetic Quantum Number (m_l), Spin Quantum Number (m_s)
- B. Orbital Quantum Number (o), Energy Quantum Number (e), Spin Quantum Number (s), Position Quantum Number (p)
- C. Principal Quantum Number (n), Energy Quantum Number (e), Spin Quantum Number (s), Angular Momentum Quantum Number (a)
- D. Spin Quantum Number (s), Orbital Quantum Number (o), Principal Quantum Number (n), Magnetic Quantum Number (m_l)

54. Which of the following is true regarding the angular momentum quantum number l ?

- A. It defines the orientation of the electron cloud
- B. It determines the energy of an electron in a given shell

- C. It indicates the shape of the orbital
 - D. It determines the spin direction of an electron
55. Which of the following correctly describes the relationship between quantum numbers?
- A. If m_l is $+1$, l must be greater than 1.
 - B. The spin quantum number must always be either $+1$ or -1 .
 - C. The value of m_l must range from $-l$ to $+l$, including zero.
 - D. n and l must be equal.
56. Which quantum number defines the energy level of an electron?
- A. Principal quantum number (n)
 - B. Azimuthal quantum number (l)
 - C. Magnetic quantum number (m_l)
 - D. Spin quantum number (m_s)
57. Which quantum number specifies the number of nodes in an orbital?
- A. Principal quantum number
 - B. Azimuthal quantum number
 - C. Magnetic quantum number
 - D. Spin quantum number
58. Which of the following statements is true?
- A. The principal quantum number determines the shape of the orbital
 - B. The azimuthal quantum number determines the energy of the orbital
 - C. The magnetic quantum number determines the orientation of the orbital
 - D. The spin quantum number specifies the angular momentum of the electron
59. The e/m value for the positive rays in the discharge tube depends upon
- A. Nature of electrode use
 - B. Composition of the gas
 - C. Nature of gas used
 - D. Pressure
60. Which of the following is the correct order of the electromagnetic spectrum in increasing order of wave number?
- A. Gamma rays, X-rays, Ultraviolet, Visible light, Infrared, Microwaves, Radio waves
 - B. Radio waves, Microwaves, Infrared, Visible light, Ultraviolet, X-rays, Gamma rays
 - C. Gamma rays, Microwaves, Infrared, Radio waves, X-rays, Ultraviolet, Visible light
 - D. Infrared, Ultraviolet, Radio waves, Gamma rays, X-rays, Microwaves, Visible light

UNIT TWO

2. CHEMICAL BONDING

Unit out come

 ***At the end of this unit, you students will be able to***

- ❖ *Differentiate intermolecular and intermolecular fore*
- ❖ *Demonstrate an understanding of the formation and general properties of substances containing ionic, covalent and metallic bond*
- ❖ *Describe the formation of ionic bonding, covalent bonding, metallic bonding*
- ❖ *Describe the formation and nature of hydrogen bonds, dipole-dipole forces and London forces*
- ❖ *Explain the bonding models: Lewis model, VSEPR model, valence bond theory, Molecular orbital theory*
- ❖ *Describing the formation of hybridization during the formation of covalent bonding*
- ❖ *Explain the relationship between Hybridization and VSEPR Theory*
- ❖ *Explain how to determine the geometry of molecules*

2.1. Introduction

2.1.1. Chemical Bonding

Chemical bonding refers to the force that holds atoms together in compounds. The stability of a molecule depends on the type and strength of these bonds.

Ionic Bonding

Definition: Ionic bonding occurs when electrons are transferred from one atom to another, resulting in oppositely charged ions that attract each other.

Property	Ionic Bond
Formation	Transfer of electrons
Participants	Metal + Non-metal
Nature	Strong electrostatic force
Example	NaCl, CaF ₂
Melting Point	High
Solubility	Soluble in polar solvents

Covalent Bonding

Definition: Covalent bonding occurs when atoms share electron pairs to achieve a stable electronic configuration.

Property	Covalent Bond
Formation	Sharing of electrons
Participants	Non-metal + Non-metal
Nature	Weak intermolecular forces
Example	H ₂ , O ₂ , CH ₄
Melting Point	Low to moderate
Solubility	Soluble in non-polar solvents

2.1.2. Intermolecular and Intramolecular Forces

✚ **Intramolecular Forces:** Strong forces within a molecule (ionic, covalent, metallic)

✚ **Intermolecular Forces:** Weaker forces between molecules (van der Waals, dipole-dipole, hydrogen bonding).

Type of Force	Strength	Examples
Hydrogen Bond	Strongest	H ₂ O, NH ₃
Dipole-Dipole	Moderate	HCl, SO ₂
Van der Waals	Weakest	CH ₄ , I ₂

2.1.3. Lewis Structure

Lewis structures represent valence electrons as dots around chemical symbols.

- **Example:** Water (H₂O)
 - Oxygen has 6 valence electrons.
 - Hydrogen shares electrons to complete octet

2.1.4. Resonance

- * Resonance occurs when a molecule has multiple valid Lewis structures.
- * Example: Benzene (C_6H_6) has alternating single and double bonds.

2.1.5. Molecular Geometry vs. Electron Set Geometry

- * Electron Set Geometry: Refers to the arrangement of electron sets (bonding pairs + lone pairs) around a central atom.
- * Molecular Geometry: Refers to the arrangement of atoms (ignoring lone pairs of electrons) in a molecule
- * The presence of lone pairs alters the molecular geometry compared to the electron set geometry.

2.1.6. Repulsion

In a molecule, electrons tend to repel each other due to their like charges. This repulsion affects the shape of the molecule and its geometry. The types of electron-electron repulsions include:

- Bonding-Bonding Repulsion: Repulsion between two bonding pairs of electrons in the valence shell of the central atom.
- Bonding-Lone Pair Repulsion: Repulsion between a bonding pair and a lone pair of electrons. Lone pairs exert more repulsion than bonding pairs because they are localized closer to the nucleus.
- Lone Pair-Lone Pair Repulsion: Repulsion between two lone pairs of electrons on the central atom.
- The strength of repulsion generally follows this order: Lone Pair–Lone Pair > Lone Pair–Bonding Pair > Bonding Pair–Bonding Pair.

2.1.7. Molecular Geometry

The following table summarizes the hybridization; electron set geometry, molecular geometry, and the role of bonding and non-bonding electrons for common types of hybridization:

Hybridization	Set of electron	Electron Set Geometry	Molecular Geometry	Bonding Electrons	Lone Pairs
sp	2	Linear	Linear	2 (bonds)	0
sp ²	3	Trigonal Planar	Trigonal Planar	3 (bonds)	0
sp ³	4	Tetrahedral	Tetrahedral	4 (bonds)	0
sp ³ (1 lone pair)	4	Tetrahedral	Trigonal Pyramidal	3 (bonds)	1
sp ³ (2 lone pairs)	4	Tetrahedral	Bent/Angular	2 (bonds)	2
sp ³ d	5	Trigonal Bipyramidal	Trigonal Bipyramidal	5 (bonds)	0
sp ³ d (1 lone pair)	5	Trigonal Bipyramidal	Seesaw	4 (bonds)	1
sp ³ d (2 lone pairs)	5	Trigonal Bipyramidal	T-shaped	3 (bonds)	2
sp ³ d ²	6	Octahedral	Octahedral	6 (bonds)	0
sp ³ d ² (1 lone pair)	6	Octahedral	Square Pyramidal	5 (bonds)	1
sp ³ d ² (2 lone pairs)	6	Octahedral	Square Planar	4 (bonds)	2

2.1.8. Effect of Bonding and Non-Bonding Electrons

Electron Type	Effect on Geometry	Location
Bonding Electrons	Determine the bonding pairs and affect molecular geometry	Shared between atoms in a bond
Non-Bonding Electrons (Lone Pairs)	Cause distortion in geometry, as they take up more space and repel bonding electrons	Found on the central atom, not involved in bonding

2.1.9. Molecular Orbital Theory (MOT)

MOT describes molecular structure using molecular orbitals formed by atomic orbitals.

- * MO theory is particularly useful for explaining more complex molecules and providing insights into properties like magnetism, bond strengths, and bond lengths.
- * When two atoms bond, their atomic orbitals (like s, p, etc.) combine to form new molecular orbitals.
 - Sigma (σ) Orbitals: These are formed from the end-on overlap of atomic orbitals (like s-s, s-p, or p_z-p_z).
 - Pi (π) Orbitals: These are formed from the side-on overlap of p-orbitals (such as p_x-p_x, p_y-p_y).

Bonding Molecular Orbitals:

- * These are lower in energy than the atomic orbitals from which they are derived.
- * Constructive overlap of atomic orbitals, leading to an increase in electron density between the nuclei, promoting and stabilizing bonding.

Antibonding Molecular Orbitals:

- * These have higher energy than the atomic orbitals from which they are derived. Electrons in antibonding orbitals tend to weaken the bond that reduced electron density between the nuclei and can destabilize the molecule.

2.1.10. MOT Vs VBT

- ❖ VB theory focuses on localized electron pairs and describes bonds using atomic orbitals.
- ❖ MO theory considers delocalized electrons across the entire molecule and treats electrons as occupying molecular orbitals that can span the whole molecule

Molecular orbitals are formed by the linear combination of atomic orbitals (LCAO), where the atomic orbitals mix to create bonding and antibonding orbitals.

Bond Order For	$BO = (N_b - N_a)/2$
If $BO > 0$	Stable molecule
If $BO = 0$	Unstable molecule
Example: O ₂	$BO = 2$ (double bond)

61. What type(s) of intermolecular forces exist between molecules of HBr and H_2S

A. Ion dipole force
B. Dipole dipole force
C. Hydrogen bonding
D. Ion induced force

🧪 Intermolecular forces (IMF) are **the attractive and repulsive forces that arise between the molecules of a substance**. These forces mediate the interactions between individual molecules of a substance. Intermolecular forces are responsible for most of the physical and properties of matter. There are different types of Interaction force between species

No	Interaction	Description	Examples
1	Hydrogen bonding	Between high electronegative molecules	H ₂ O, NH ₃ , HF, CH ₃ CH ₂ OH
2	Dipole-dipole force	Between polar molecules	HBr, HI, H ₂ S, HCl
3	London force	Between non polar molecules	CH ₄ , CO ₂ , CCl ₄ , C ₆ H ₆
4	Ion dipole force	Between ion and polar molecules	Fe ²⁺ in SO ₂
5	Ion-induce dipole force	Between ion and non-polar molecules	Fe ²⁺ in Hemoglobin
6	Dipole induced dipole force	Between polar and non-polar molecules	HBr in CH ₄

- ✓ To identify intermolecular forces, it is useful to classify the species being considered as (1) non- polar molecules, (2) polar molecules, and (3) ions.
- ✓ Keep in mind that dispersion forces exist between all species.
- ✓ Due to the large electronegativity difference between hydrogen and bromine/sulfur, the H–Br bond and H–S bond are polar. The molecular interactions, thus, are dipole-dipole interactions.

A. P – Cl B. S – O C. Si – F D. C – N

Given the electronegativity data of elements:
N (3.0), Cl (3.0), C (2.5), H (2.2), Si (1.8), P (2.1), F (4.0), S(2.58), O(3.44)

Calculate the Electronegativity difference, The more electronegativity difference the more polar.

- P-Cl..... $3.0-2.1=0.9$
- S-O..... $3.44-2.58=0.86$
- Si-F..... $4.0-1.8=2.2$ The most polar
- C-N..... $3.0-2.5=0.5$ The least polar
- Generally: In polarity, $\text{Si-F} > \text{P-Cl} > \text{S-O} > \text{C-N}$

63. Hybridization of sp^3d^2 orbital is associated with structure.

A. Trigonal bipyramidal

B. Octahedral

C. Tetrahedral

D. planar

Brief:

Set of electron	Designation	No of Bonding	Repulsion	Electron set Geomet (ideal)	Molecular Geomet (True)	Hybridization	Bond angle	Examples
2	AX_2	2+0	Bp-BP	Linear	Linear	sp	180°	$BeCl_2$, CO_2 , HCN , I_2
3	AX_3	3+0	Bp-BP	Trigonal Planar	Trigonal Planar	sp^2	120°	BF_3 , SO_3 , $AlCl_3$, NO_3^-
	AX_2E	2+1	Bp - BP Bp - LP	Trigonal Planar	Bent, V-shape	sp^2	$<120^\circ$	SO_2 , O_3^{2-} , NO_2^- , XeO_3
4	AX_4	4+0	Bp - BP	Tetrahedral	Tetrahedral	sp^3	109.5°	CH_4 , NH_4^+ , CCl_4 , $SiCl_4$
	AX_3E	3+1	Bp -BP Bp -LP	Tetrahedral	Trigonal Pyramidal	sp^3	107°	NH_3 , PCl_3 , $AsCl_3$, $SbCl_3$
	AX_2E_2	2+2	Bp-BP Bp-LP Lp-LP	Tetrahedral	Bent, V-shape	sp^3	$<109.5^\circ$	H_2O , H_2S , $SnCl_2$, OF_2
5	AX_5	5+0	Bp-BP	Trigonal Bipyramidal	Trigonal Bipyramidal	sp^3d	90° , 120°	PCl_5 , AsF_5 , VCl_5 , SbF_5
	AX_4E	4+1	Bp-BP Bp-LP	Trigonal Bipyramidal	See-saw	sp^3d	$<90^\circ$, $<120^\circ$	SF_4 , $SeCl_4$, $XeCl_4$, ICl_4^-
	AX_3E_2	3+2	Bp-BP Bp-LP Lp-LP	Trigonal Bipyramidal	T-shaped	sp^3d	$<90^\circ$	ClF_3 , BrF_3 , ICl_3 , $BrCl_3$
	AX_2E_3	2+3	Bp-BP Bp-LP Lp-LP	Trigonal Bipyramidal	Linear	sp^3d	180°	XeF_2 , I_3^- , $XeCl_2$, I_2Cl_2
6	AX_6	6+0	Bp-BP	Octahedral	Octahedral	sp^3d^2	90°	SF_6 , $MoCl_6$, $ReCl_6$
	AX_5E	5+1	Bp-BP Bp-LP	Octahedral	Square Pyramidal	sp^3d^2	$<90^\circ$	BrF_5 , IO_5 , PF_5 , SbF_5
	AX_4E_2	4+2	Bp-BP Bp-LP Lp-LP	Octahedral	Square Planar	sp^3d^2	90°	XeF_4 , $PtCl_4^{2-}$, $PdCl_4^{2-}$, $IrCl_4^-$

64. According to the molecular orbital theory the bond order of O_2^- is

- A. 2.5 B. 1.5 C. 2 D. 0.5

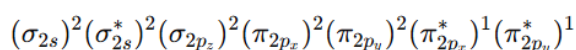
Brief:

According to **Molecular Orbital Theory (MOT)**, the bond order of a molecule is given by the formula:

$$\text{Bond Order} = \frac{(\text{Number of bonding electrons} - \text{Number of antibonding electrons})}{2}$$

For O_2^- (superoxide ion):

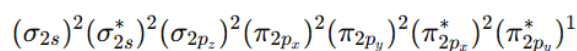
1. The electronic configuration of neutral O_2 (16 electrons per oxygen, total 32 electrons) in molecular orbitals is:



2. O_2^- has **one extra electron**, making it 33 electrons in total.

- The extra electron goes into the π (antibonding) orbitals*.

- New configuration:



3. Count bonding and antibonding electrons:

- **Bonding electrons: 10** (σ and π bonding orbitals)
- **Antibonding electrons: 7** (σ^* and π^* antibonding orbitals)

65. Which of the following compounds contain both ionic and covalent bonds?

- A. CO_2 B. Na_2CO_3 C. Cl_2O D. $BaCl_2$

Brief:

🚦 Ionic bonding

- It is electrovalent bond
- Formed between cation and anion by transferring of electron
- Formed between metal (low IE) and non-metal (high EA)

🚦 Covalent bonding

- Formed between non metal
- Formed by sharing of electron (Equal sharing (Non polar) or Unequal sharing (polar))

🚦 There are some compounds formed by both ionic and covalent bond like:



66. Which of the following is the most reasonable Lewis formula for FCN?

- A. $F=C=N$ B. $F=C-N$ C. $F-C \equiv N$ D. $F-C-N$

Brief: I. Determine the total number of valence electrons.

$$\text{Total v.e} = F(7) + C(4) + N(5) = 16e^-$$

II. Write the skeleton structure of the molecule.

*.Hydrogen atoms are always terminal.

*.Central atoms are generally those with the lowest EN.

*.Carbon atoms are always central

III. Use two valence electrons to form each bond in the skeleton structure.

IV. Try to satisfy the octets of the atoms by distributing the remaining valence electrons as nonbonding electrons.

V. If the central atom is not still satisfied octet rule, make further multiple bond for which electrons are contributed from terminal atom (=coordinative covalent bond).

- Lewis structure of FCN is: $\text{F}-\text{C}\equiv\text{N}$

The lewis structure of molecules may consist of

- * Bonding electron
- * Non-bonding electron (lone pair)

Bonding pair = $\frac{1}{2}$ bonding electron

Lone pair = $\frac{1}{2}$ non bonding electron

67. Which molecule is linear? A. H_2O B. NH_3 C. CO_2 D. BF_3

Brief:

✚ Write the Lewis structure of each molecule

✚ State their designation and bond respectively

molecules	Designation	Bond Angle
CF_4	AX4	
SF_6	AX6	
CO_2	AX2	
BF_3	AX3	

68. Which of the following compounds contains only covalent bonds?

KCl, PH_3 , O_2 , B_2H_6 , H_2S

A. KCl, B_2H_6 , PH_3 B. O_2 , PH_3 , B_2H_6 C. KCl, H_2S D. KCl, B_2H_6

Covalent molecules are produced between non-metals by sharing of electron

❖ Equally,

Provides non-polar bond and the molecule is also non-polar

❖ Unequally,

Provides polar bond and the molecule may be polar or non-polar based on the dipole Moment

If dipole moment = 0, non-polar

If dipole moment $\neq$ 0, polar

Examples: PH_3 , O_2 , H_2S , O_2 , H_2S , B_2H_6

69. Which diatomic species has the greatest bond energy?
 A. NO B. NO^+ C. O_2 D. O_2^{+2}

Brief:

- ✚ The strongest bond for a diatomic species is that of **carbon monoxide**, CO
- ✚ The strongest bond for a mononuclear diatomic species is that of nitrogen, N_2
- ✚ When a bond is strong, there is higher bond energy because it takes more energy to break a strong bond.
- ✚ The higher bond order, the shorter bond length and, the **the greater the bond energy**.

1. NO (Nitric Oxide)

- Total electrons = 15
- MO configuration: $\sigma(2s)^2\sigma^*(2s)^2\pi(2p)^4\sigma(2p)^2\pi^*(2p)^1$
- Bond order = $(8 - 3) / 2 = 2.5$

2. NO^+ (Nitrosonium Ion)

- Total electrons = 14
- MO configuration: $\sigma(2s)^2\sigma^*(2s)^2\pi(2p)^4\sigma(2p)^2\pi^*(2p)^0$
- Bond order = $(8 - 2) / 2 = 3$

3. O_2 (Oxygen Molecule)

- Total electrons = 16
- MO configuration: $\sigma(2s)^2\sigma^*(2s)^2\sigma(2p)^2\pi(2p)^4\pi^*(2p)^2$
- Bond order = $(8 - 4) / 2 = 2$

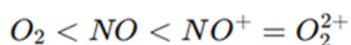
4. O_2^{2+} (Diperoxide Ion)

- Total electrons = 14
- MO configuration: $\sigma(2s)^2\sigma^*(2s)^2\sigma(2p)^2\pi(2p)^4\pi^*(2p)^0$
- Bond order = $(8 - 2) / 2 = 3$

Comparing bond orders:

- $\text{O}_2 \rightarrow 2$
- $\text{NO} \rightarrow 2.5$
- $\text{NO}^+ \rightarrow 3$
- $\text{O}_2^{2+} \rightarrow 3$

Thus, the increasing order of bond order is:



70. The structure of IF_5 is based on which configuration?

- A. Trigonal bipyramidal B. Tetrahedral C. octahedral D. square bipyramidal

Brief:

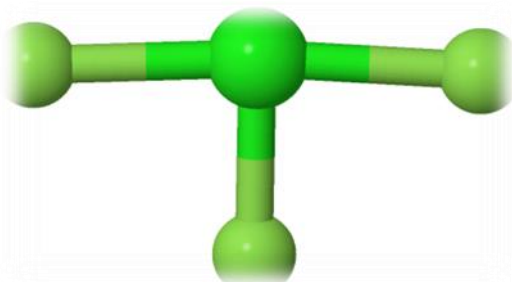
- ✚ Draw the lewis structure of the molecule IF_5
- ✚ It is AX_5E type
- ✚ Any molecules having AX_5E types are square pyramidal

71. What is the shape of ClF_3 , BrF_3 and IF_3

- A. Square pyramidal B. Square planar C. Sea saw D. T-shape

Brief:

All of the given molecules (ClF_3 , BrF_3 and IF_3) are AX_3E_2 T-shape structure.

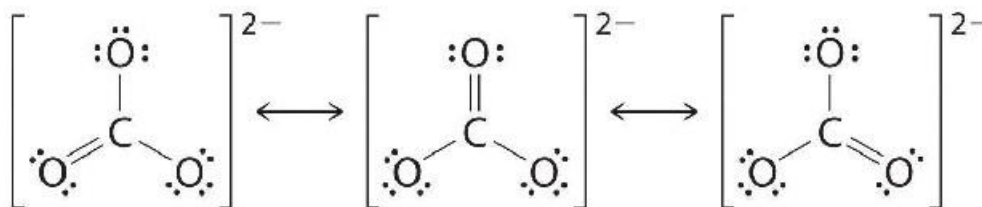


72. As experimentally verified, the carbonate ion, CO_3^{2-} , contain one double bond. How many resonance structures are possible in the representation of its Lewis structures?

- A. 2 B. 4 C. 3 D. 5

Brief:

- There are three resonance structure of CO_3^{2-}



73. What is the hybridization of sulphur atom in the SF_5^- ion?

- A. SP^3d^2 B. SP^3d C. SP^3 D. SP^2

Brief:

- ✓ The number of hybrid orbital = Number of set of electron
- ✓ The geometry designation of SF_5^- is AX_5E having 6 set of electron
- ✓ The hybridization scheme having 6 hybrid orbital is SP^3d^2

74. How many sigma(δ) and Pi(π) bonds does the molecule $\text{CH}_3\text{CHCHCOOH}$ contain?
A. 9 δ and 2 π B. 9 δ and 4 π C. 11 δ and 2 π D. 13 δ and 4 π

Brief:

After writing the detailed structure of the compound (CH₃-CH=CH-CH-COOH)

The number of sigma(δ) and Pi(π) bonds are 11 and 2 respectively.

75. Given the species: H_2O , NH_2^- , Cl_2O . Which of these does **NOT** represent these species in common?
- A. sp^3 hybridization C. Tetrahedral electron pair arrangement
B. Covalent bond D. None of these.

Brief

* H_2O (Water)

- ✓ Oxygen has sp^3 hybridization.
- ✓ There are two lone pairs and two bonded pairs on oxygen.
- ✓ The electron pair geometry is tetrahedral.
- ✓ The bond type is covalent.

* NH_2^- (Amide ion)

- ✓ Nitrogen has sp^3 hybridization.
- ✓ There are one lone pair and two bonded pairs on nitrogen.
- ✓ The electron pair geometry is tetrahedral.
- ✓ The bond type is covalent.

* Cl_2O (Dichlorine monoxide)

- ✓ Oxygen has sp^3 hybridization.
- ✓ There are two lone pairs and two bonded pairs on oxygen.
- ✓ The electron pair geometry is tetrahedral.
- ✓ The bond type is covalent.

- sp^3 Hybridization → Common to all three

-Covalent Bond → Common to all three

-Tetrahedral Electron Pair Arrangement → Common to all three

76. The carbon atoms in ethane (C_2H_6), ethene (C_2H_4) and ethyne (C_2H_2) Provide, respectively, examples of the three common types of hybridization corresponding to:

- A. sp, sp^2, sp^3 B. sp, sp^3, sp^2 C. sp^3, sp^2, sp D. sp^3, sp, sp^2

Brief:

- ✚ Single bond Only sigma bond like ethane ethane (C_2H_6)..... sp^3 hybridization
- ✚ Double bond.....One sigma + one pi bond like ethene (C_2H_4)..... sp^2 hybridization
- ✚ Triple bond.....One sigma + two pi bond like ethyne (C_2H_2)..... sp hybridization

77. Which of the following compounds has dipole moment zero?

- A. CH_4 B. H_2O C. CH_3Cl D. NH_3

Brief:

- ✚ The polarity of polyatomic molecules is depends on the dipole moment
- ✚ Dipole moment =non polar molecule while dipole moment $\neq 0$ for polar molecules
- ✚ Dipole moment = 0, Iff the bond moments are equal in magnitude and opposite direction
- ✚ Dipole moment $\neq 0$ for unequal magnitude and/or/ the same direction of bond moment.
 - CH_4 Equal magnitude + opposite direction non-polar
 - H_2O Equal magnitude + the same direction polar
 - CH_3Cl unequal magnitude + different direction polar
 - NH_3 Equal magnitude + the same direction non-polar

78. Which of the following correctly describes a π bond?

- A. It is formed by the end-to-end interaction of p -orbitals and has a low Electron density on the internuclear axis.
- B. It is formed by the parallel interaction of p -orbitals and has a high electron density just above and below the internuclear axis.
- C. It is formed by the interaction of s -orbitals and has a low electron density on the antinuclear axis.
- D. It is formed by the interaction of s -orbitals and has a high electron density on the internuclear axis

Brief:

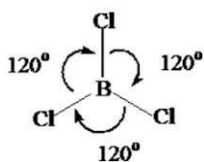
- ✚ A π bond is formed by the side-to-side overlap of two p orbitals. It is a type of covalent bond that occurs when two atoms share electrons through the overlap of two p orbitals.

79. Which of the following molecules would you expect to be planar?

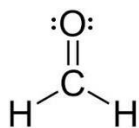
- A. CH_4 B. XeF_4 C. BCl_3 D. HCHO

Brief:

- XeF_4 is a **non-polar chemical**. Because XeF_4 has a symmetrical geometrical structure, it is square planar. The dipoles across the Xe-F bond cancel each other out, resulting in a zero net dipole.
- The C-H bonds in **methane (CH_4)** are polar. However, a molecule of methane is non-polar. Specifically, the dipole moment of methane is zero.
- BCl_3 is a nonpolar molecule. A polar bond requires an asymmetrical shape shift in electron density to generate an electrical dipole, which BCl_3 does not have



- Formaldehyde (CH_2O) is a polar molecule due to the unequal distribution of electron density. The electronegativity difference between the hydrogen and carbon atoms is negligible, but the difference between the carbon and oxygen atoms is significant enough to induce polarity.



80. Which of the following species has the smallest H-X-H bond angle where X is the central atoms?

- A. OH_2 B. NH_3 C. CH_4 D. BH_3

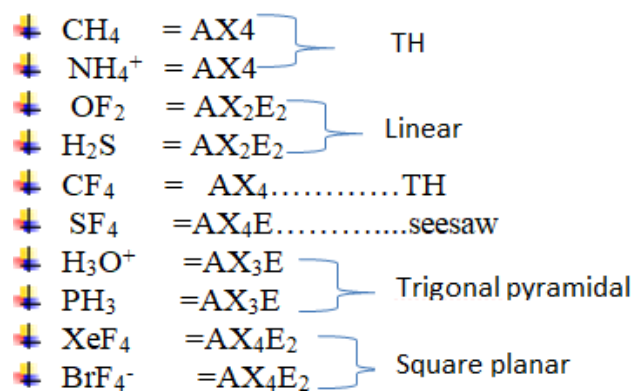
Brief:

- $\text{OH}_2 = \text{AX}_2$ (H-X-H)
- $\text{CH}_4 = \text{AX}_4$
- $\text{NH}_3 = \text{AX}_3\text{E}$
- $\text{BH}_3 = \text{AX}_3$

81. The species in which of the following pairs have decidedly different geometries?

- A. CH_4 & NH_4^+ C. OF_2 & H_2S E. CF_4 & SF_4 B. H_3O^+ & PH_3 D. XeF_4 & BrF_4^-

Brief:



82. Which of the following ionic compounds has the smallest lattice energy?

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

Brief

- Lattice energy is a measure of the strength of the ionic bonds in an ionic compound.
- It provides insight into several properties of ionic solids including their volatility, their solubility, and their hardness.
- The lattice energy of an ionic solid cannot be measured directly.
- The factors affecting lattice energy are size of ion and distance between
- As the magnitude of charge increase (size decrease) lattice energy increase while solubility decrease
- Down the group, the lattice energy decreases. This is due to the fact that with the increase in atomic size, the distance between their nuclei increases. Therefore, decreasing the force of attraction and increasing the inter-electronic repulsion between them.
- The highest lattice energy is present in **CsF** because F is most electronegative among halogens so the force of electrostatic attraction is high between Cs and F.
- Lattice energy: CsF < RbF < KF < NaF < LiF
CsI < NaF < LiI < LiF
NaF < Nil < NaBr < NaI
- Solubility is the reciprocal of lattice energy.
- Both hydration energy and lattice energy are decrease down a group, but

In metal hydroxide hydration energy decrease more rapidly than lattice energy does.

83. Of the following the one with the largest bond angle is

- A. BF_3 B. CCl_4 C. H_2S D. H_3O^+

Brief:

- $\text{BF}_3 \dots \text{AX}_3 \dots 120^\circ$
- $\text{CCl}_4 \dots \text{AX}_4 \dots 109.5^\circ$
- $\text{H}_2\text{S} \dots \text{AX}_2\text{E} \dots 107^\circ$
- $\text{H}_3\text{O}^+ \dots \text{AX}_3 \dots 109.5^\circ$

84. Choose the wrong combination arrangement (self-practice)

<u>Molecule/ion</u>	<u>designation- Type</u>	<u>Hybridization</u>
A. ClF_3	AX_3E_2	sp^3d
B. H_2O	AX_3E_2	sp^3
C. CO_2	AX_2	sp
D. OF_2	AX_2	sp

85. Number of σ and π bond respectively in $\text{CH} \equiv \text{C} - \text{CH}_2 - \text{CH} = \text{CH}_2$ is ((self-practice)

- A. 1&2 B. 9&4 C. 10 &3 D. 8 &

86. Which one of the following molecules /molecular ions is diamagnetic according to the molecular orbital theory? A. NO^+ B. O_2 C. F_2^+ D. C_2

Brief:

- ✚ If all electrons are paired then diamagnetic
- ✚ If electrons are unpaired then paramagnetic
- ✚ Nitric oxide is paramagnetic because of the presence of an odd electron (unpaired electron).
- ✚ NO^+ has no unpaired electron so it is **diamagnetic** in nature
- ✚ O_2 has unpaired electron those are **paramagnetic**
- ✚ F_2^+ has only 1 unpaired electron. This means it is **paramagnetic**.
- ✚ C_2 has no unpaired electron, so it is **diamagnetic** in nature.

87. In order to calculate ΔH_f of KCl by using Born Haber cycle, which of the Following parameter is not required?

- A. ΔH fusion of K C. ΔH_{LE} (KCl)
 B. ΔH sublimation K. D. Electron gain enthalpy of K

Brief:

- Steps required in Born Haber process
 - ✚ Atomization (enthalpy of sublimation)....
 - ✚ Dissociation (enthalpy of Dissociation)

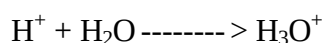
- ✚ Ionization (Enthalpy of ionization energy)
- ✚ Adding of electron to Cl (Enthalpy of Electron Affinity)
- ✚ Ionic interaction (Lattice energy)

88. When H^+ forms a bond with H_2O to form the hydronium ion H_3O^+ , this bond is called a coordinate covalent bond because ____.

- A. both bonding electrons come from the oxygen atom
- B. it forms an especially strong bond
- C. the electrons are equally shared
- D. the oxygen no longer has eight valence electrons

Brief:-

- ✚ In coordinative covalent bond, both bonding electrons are contributed from one of the two electrons.
- ✚ When H^+ forms a bond with H_2O to form hydronium ion, H_3O^+ , this bond is called a coordinate covalent bond because a. both bonding electrons come from the oxygen atom.



89. Of the possible bonds between carbon atoms (single, double, and triple), ____.

- A. A triple (sigma + 2pi) bond is longer than a single bond
- B. A double (sigma + pi) bond is stronger than a triple bond
- C. A single (sigma) bond is stronger than a triple bond
- D. A double (sigma + pi) bond is longer than a triple bond
- E. A single (sigma) bond is stronger than a double bond

Brief:

- ✚ The possible bonds between carbon atoms
 - Single bond (-).....1-sigma ONLY
 - Double bond (=).....1-Sigma bond + 1-Pi bond
 - and triple bond ($\equiv$).....1-sigma bond + 2-pi bond
 - Strength:- triple bond > double bond > single bond
 - Length:- single bond > double bond > triple bond
 - Energy:- triple bond > double bond > single
 - Reactivity:- triple bond > double bond > single
 - Stability:- single bond > double bond > triple bond

90. Of the bonds $\text{C} - \text{N}$, $\text{C} = \text{N}$, $\text{C} \equiv \text{N}$ the $\text{C} - \text{N}$ bond is _____(self-practice)

- A. strongest/shorte B. strongest/longest Weakest/shortest D. W e a k e s t /longest

91. Which of the ions is the most abundant in sea water

- A. Na^+ B. Ca^{+2} C. Cl^- D. HCO_3^-

Brief:

✚ Chlorine, present as chloride ion, is the most abundant ion and makes up about 55 per cent by weight of the dissolved material.

✚ It is rarely measured except in combination with other halides in the chlorinated determination

Brief:

92. The basis of the VSEPR model of molecular bonding is _____

- A. regions of electron density on an atom will organize themselves so as to maximize s- character
- B. regions of electron density in the valence shell of an atom will arrange themselves so as to maximize overlap
- C. atomic orbitals of the bonding atoms must overlap to form a bond
- D. to form electron domains in the valence shell of an atom will arrange themselves so as to minimize repulsions
- E. hybrid orbitals will form as necessary to, as closely as possible, achieve spherical symmetry

Brief:

✚ According to VSEPR theory, electron domains of any atom, whether bonded or unbounded, arrange themselves in some specified direction to minimize repulsions.

✚ VSEPR theory assumes that each atom in a molecule will achieve a geometry that minimizes the repulsion between electrons in the valence shell of that atom.

✚ The premise of the VSEPR theory is that electron pairs located in bonds and lone pairs repel each other and will therefore adopt the geometry that places electron pairs as far apart from each other as possible.

✚ This theory is very simplistic and does not account for the subtleties of orbital interactions that influence molecular shapes; however, the simple VSEPR counting procedure accurately predicts the three-dimensional structures of a large number of compounds, which cannot be predicted using the Lewis electron-pair approach.

93. The heat formation of KF(s) is -563KJ/mol, the ionization energy of k(g) is 419KJ/mol and the heat of sublimation of Potassium is 88KJ/mol the electron affinity of F(g) is -322KJ/mol and the F-F bond energy is 158KJ/mol . Calculate the lattice energy of KF(s) A.-1036KJ/mol B. -563KJ/mol C. -982KJ/mol D.-739KJ.mol e.none.

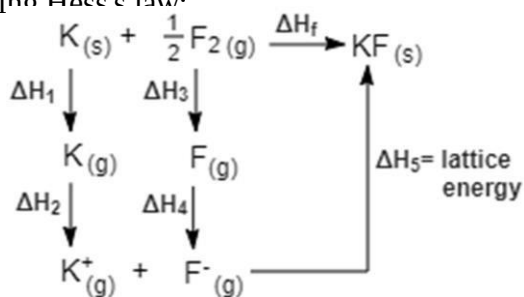
Brief:

➤ Information:

- The ΔH_f of KF(s) is -563 kJ/mol.
- The ionization energy of potassium K(g) is 419 kJ/mol
- Enthalpy of sublimation of potassium is 88 kJ/mol.
- The electron affinity of F(g) is -322 kJ/mol
- F-F bond enthalpy is 158 kJ/mol.
- Lattice energy (in kJ/mol) of KF(s) =?

✚ Born-Haber cycle for the formation of KF is:

Using Hess's law:



- $\Delta H_f = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4 + \Delta H_5$
- $-563 = 88 + 419 + 79 - 322 + \Delta H_5$
- $\Delta H_5 = -827 \text{ kJ/mol}$

94. Which of the following has the bonds correctly arranged in order of increasing polarity?

- A. Be — F, Mg — F, N — F, O — F
- B. O — F, N — F, Be — F, Mg — F
- C. O — F, Be — F, Mg — F, N — F
- D. N — F, Be — F, Mg — F, O — F
- E. Mg — F, Be — F, N — F, O — F

Brief:

- ✚ State the electron negativity of each elements from different source
- ✚ Calculate the electron negativity difference between each pair of atoms
- ✚ Give the order from the less polar (smallest electron negativity difference) to most polar (highest electron negativity).

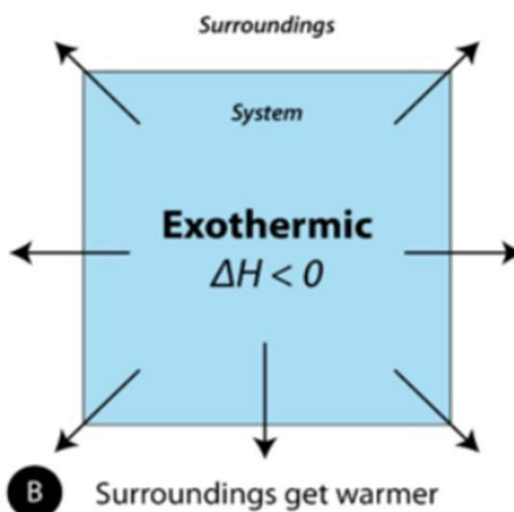
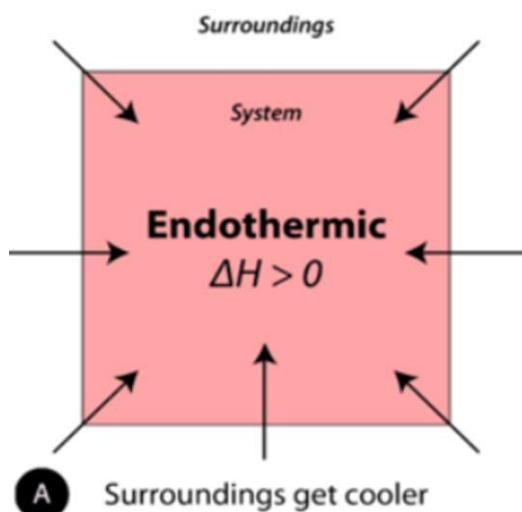
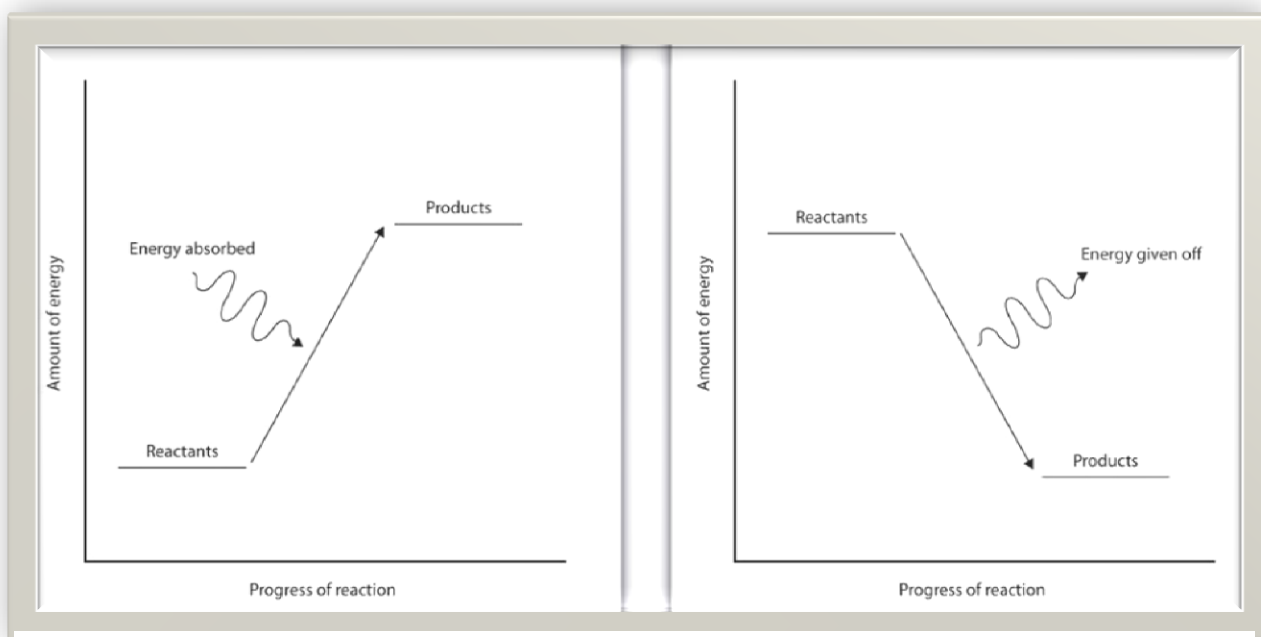
95. Which of the following pair is an exothermic process?

- A. Heat of sublimation and ionization energy B. Heat of formation and electron affinity
B. Ionization energy and bond formation D. Heat of formation and bond dissociation.

Brief:

Endothermic reaction

exothermic reaction



Physical process	Process	Enthalpy change	Exothermic /Endothermic
Solid to Liquid	Melting	Enthalpy of fusion	Endothermic ($\Delta H_{\text{fusion}} = +ve$)
Liquid to Gas	Vaporization /Boiling	Enthalpy of Vaporization.	Endothermic ($\Delta H_{\text{vap}} = +ve$)
Solid to Gases	Sublimation	Enthalpy of Sublimation	Endothermic ($\Delta H_{\text{subl}} = +ve$)
Liquid to Solid	Freezing /crystallization	Enthalpy of crystallization	Exothermic ($\Delta H_{\text{cryst}} = -ve$)
Gases to Liquid	Condensation	Enthalpy of Condensation	Exothermic ($\Delta H_{\text{cond}} = -ve$)
Gases to solid	Deposition	Enthalpy of deposition	Exothermic ($\Delta H_{\text{dep}} = -ve$)

96. The electron-domain geometry of the central atom in OF_2 is _____while the molecular geometry would be described as_____.

- A. linear, bent B. trigonal planar, bent C. tetrahedral, bent
 B. trigonal bipyramidal, linear D. octahedral, square pyramidal

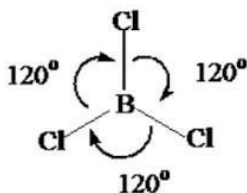
Brief:

97. The molecular geometry of the BCl_3 molecule is_____and this molecule is_____

- A. trigonal pyramidal, nonpolar
 B. trigonal planar, polar
 C. trigonal planar, nonpolar
 D. trigonal bipyramidal, polar

Brief:

- ✚ BCl_3 is a nonpolar molecule. A polar bond requires an asymmetrical shape shift in electron density to generate an electrical dipole, which BCl_3 does not have.
- ✚ The chemical molecule BCl_3 has an sp^2 hybridization type.
- ✚ Thus, there exists no polarization of charges across the BCl_3 patch.
- ✚ BCl_3 , boron has only 6 electrons in it's valence shell (3 from boron and 3 from chlorine atoms) and, one vacant p-orbital. BCl_3 does not obey octet rule.



98. Which one of the following molecules is not a multiple bond?

- A. O_2 B. Cl_2 C. N_3 D. CO_2

Brief:

valence Electrons	Share 1 pair (2 electrons)	Share 2 pairs (4 electrons)	Share 3 pairs (6 electrons)
Bond Length	Longest	Intermediate	Shortest
Bond Strength	Weakest	Intermediate	Strongest
Reactivity	Lowest	Intermediate	Highest
Rotation Around Bond	Yes	No	No
Orbitals	One sigma	One sigma, one pi	One sigma, two pi
Notation	Single dash (C-C)	Double dash (C=C)	Triple dash (C≡C)
Examples	Alkanes, H_2 , X_2 , HX (X=halogens)	O_2 , CO_2 , Alkenes, SO_2 , SO_3	N_2 , CO, Alkynes

99. Which of the following molecules has the highest bond strength? (self-practice)

- A. C_2H_2 B. C_2H_4 C. CO_2 D. H_2S

100. Which one of the following molecules /molecular ions is paramagnetic according to the molecular orbital theory?

- A. O_2^{-2} B. O_2 C. F_2 D. O_2^{+2}

Brief:

- Materials with unpaired electrons are paramagnetic and attracted to a magnetic field, while those with all-paired electrons are diamagnetic and repelled by a magnetic field

🚦 Diamagnetic

- All electrons are **paired**

- If total no. of electrons in the molecule are 10, 16 or odd electrons, the molecule is paramagnetic.

🚦 Paramagnetic

- All electrons are **not paired**

- If total no. of electrons are even except 10 & 16 then the molecule is diamagnetic

Category	Paramagnetic	Diamagnetic
Definition	Substance gets weakly magnetized on application of a magnetic field.	Substances oppose the magnetism under the applied magnetic field.
Permanent dipole moment	Yes	No
Direction	Same direction as of the applied Magnetic field.	Opposite
Electron arrangement	There is unpaired electron	All electrons are paired.
Effect of temperature	Yes	No
Examples	O_2	N_2

Then Write Electron configuration of: O_2^{-2} , O_2 , F_2 , O^{+2}

101. What is the hybridization of phosphorus atom in PCl_5

- A. SP^3d B. Sp^3 C. Sp^2 D. SP^3d^2

Brief:

✚ Hybridization is the process in which two complementary single-stranded DNA and/or RNA molecules bond together to form a double-stranded molecule.

✚ No of hybrid orbitals = No of set of electron H_3O^+

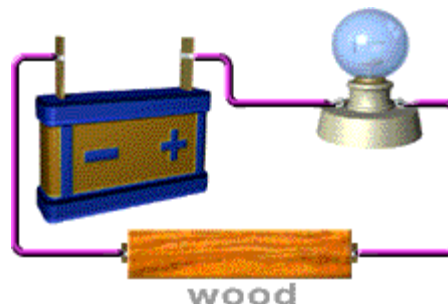
✚ The Lewis structure of H_3O^+ $\left[\begin{array}{c} \text{H}-\ddot{\text{O}}-\text{H} \\ | \\ \text{H} \end{array} \right]^+$

102. Which one of the following substance is a non –conductor of electricity?

- A. Graphite B. $\text{MgCl}_2(\text{s})$ C. Silver(s) D. $\text{H}_2\text{SO}_4(\text{aq})$

- **Conductivity** is the measure of the ease at which an electric charge or heat can pass through a material. A conductor is a material which gives very little resistance to the flow of an electric current or thermal energy. Materials are classified as metals, semiconductors, and insulators. Metals are the most conductive and insulators (ceramics, wood, plastics) the least conductive.

Electrical conductivity tells us how well a material will allow electricity to travel through it. Many people think of copper wires as something that has great electrical conductivity.



Thermal conductivity is the ease upon which thermal energy (heat for most purposes) can move through a material. Some materials like metals allow heat to travel through them quite quickly. Imagine that with one hand you are touching a piece of metal and with the other, a piece of wood. Which material would feel colder? If you said, "metal," you would be correct. But, in fact, both materials are in fact the same temperature. This is relative thermal conductivity. Metal has a higher heat transferability, or thermal conductivity, than wood, letting the heat from your hand leave faster. If you want to keep something cold the best idea is to wrap it in something that does not have a high heat transferability, or high thermal conductivity, this would be an insulator. Ceramics, and polymers are usually good insulators, but you have to remember that polymers usually have a very low melting temperature. That means if you are designing something that will get very hot the polymer might melt, depending on its melting temperature.

✚ Electrical and thermal conductivity are closely related. For the most part good electrical conductors are also good thermal conductors.

✚ Many products will contain both conductors and insulators- the conductors take the electricity or thermal energy where it is wanted and the insulators prevent it from getting where it isn't wanted.

➤ Generally, There are two kinds of conductivity

i. Metallic/electronic conductivity

- Conduct electric current due to delocalized electron
- Example; metals (like Ag, Cu, Cr) and Graphite
- The conductivity order of some commonly used metals and alloys, based on equal sizes are:

~Pure silver > Pure copper > Pure gold Aluminum > Zinc Nickel > Brass Bronze > Iron > Platinum Steel > Lead > Stainless steel

- Graphite is the only non-metal that conduct electricity
- Graphite conducts electricity whereas diamond does not because in diamond the carbon atoms are bonded to other carbon atoms and all the valence electrons are bonded. Whereas in graphite each carbon atom is only bonded to three other atoms. Therefore, there are delocalized (free) electrons in graphite due to 1 free valence electron in each carbon atom, which can move and carry a charge so graphite conducts electricity.

ii. Electrolytic/ionic conductivity

- Conductivity of materials in aqueous /molten state
- It is due to providing ions (electric current carriers)
- Eg. Soluble salt, strong acid, strong base

NB: Substances those are conducting electric current in aqueous or molten state are electrolyte

There are three form of electrolyte: Strong electrolyte, weak electrolyte, non-electrolyte

Strong Electrolytes	strong acids	HCl, HBr, HI, HNO ₃ , HClO ₃ , HClO ₄ , and H ₂ SO ₄
	strong bases	NaOH, KOH, LiOH, Ba(OH) ₂ , and Ca(OH) ₂
	Soluble salts	NaCl, KBr, MgCl ₂ ,
Weak Electrolytes	weak acids	Organic acids , H ₂ CO ₃ , H ₃ PO ₄
Non electrolyte	Most covalent compounds/molecules	Glucose, ammonia, pure water,

103. Which of the following element has the highest melting point?

- A. Iodine B. Tungsten C. Mercury D. Bromine

Brief

As a pure element **tungsten** has the highest melting point, about 3500 C

The top 10 metals in their melting point are

Tungsten (W) > Rhenium (Re) > Tantalum (Ta) Molybdenum (Mo) > Niobium (Nb) > Iridium (Ir) > Hafnium (Hf) > Rhodium (Rh) > Chromium (Cr) > Titanium (Ti)

104. Given the molecules :

CH₄, NH₃, H₂S, CH₄, CH₃OH, CCl₄, OF₂, NF₃, H₂O, CO₂, BCl₃, PCl₃, O₂, HF, HBr,

Which one of the following set of molecules correctly described correct intermolecular force of attraction?

Options	Set of molecules	Intermolecular force
A	CH ₄ , CH ₄ , CCl ₄ , O ₂ , BCl ₃ , CO ₂	London force
B	NH ₃ , CH ₃ OH, H ₂ O, HF, NF ₃	Hydrogen bonding
C	H ₂ S, CCl ₄ , NF ₃ , PCl ₃ , HBr, HF	Dipole - dipole force
D	H ₂ S, NF ₃ , CCl ₄ , NF ₃ , PCl ₃ , HBr	Dipole - dipole force

Brief:

✚ Molecules having London forces are: CH₄, CH₄, CCl₄, O₂, CO₂

✚ Molecules having dipole-dipole forces: H₂S, NF₃, PCl₃, HBr, OF₂

✚ Molecules having Hydrogen bonding: NH₃, CH₃OH, H₂O, HF

105. Which species has the largest F-A-F bond angle where A is the central atom? (self-practice)

- A. BF₃ B. CF₄ C. NF₃ D. OF₂

106. Crystalline solids

- A. have their particles arranged randomly B. are usually very soft
C. has high ordered structure D. has low melting point

Brief:

✚ Crystalline solids have a very ordered, repeating structure. This repeating structure is what gives them their characteristic shape and makes them very strong and difficult to break.

✚ The atoms in a crystalline solid are arranged in a three-dimensional, geometric pattern called a crystal lattice

✚ The melting point is the temperature at which the solid turns into a liquid.

✚ The higher the melting point, the more difficult it is to melt the solid.

✚ Crystalline solids generally have very high melting points because of their strong, ordered structure.

- ✚ This makes them useful for applications where high temperatures are involved, such as in metallurgy and ceramic manufacturing.

107. What is the shape of ClF_3 , BrF_3 and ICF_3 ?

- A. Square pyramidal B. Square planar C. Sea saw D. T-shape

Brief:

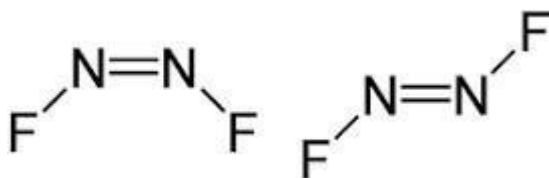
✚ The molecular designation of ClF_3 , BrF_3 and ICF_3 is AX_3E_2

✚ The molecular geometry of all of these molecules is T shape.

108. N_2F_2 exists in cis and Trans forms a bonding model that can account for observation described the bond between the two nitrogen atoms as

- A. One σ bond only C. One σ bond and one π bond only
B. One π bond only D. One σ bond and two π bonds

Brief:



Consist of 3 σ bond and 1 π bond

The bond between nitrogen atoms are: σ and 1 π

109. Which resonance structure is the most stable?

- A. The one with the most formal charges C. The one with the highest energy
B. The one with formal charges close to zero D. The one with the most lone pairs

Brief:

When evaluating resonance structures, the most stable structure generally follows these guidelines:

Formal charges should be minimized – The best resonance structure is the one where most atoms have formal charges as close to zero as possible.

Electronegativity considerations – If formal charges are present, negative charges should be on the most electronegative atoms (e.g., oxygen, fluorine), while positive charges should be on less electronegative atoms (e.g., carbon).

1. Octet rule satisfaction – Structures where all atoms satisfy the octet rule (except hydrogen) are more stable.
2. Delocalization – Structures that allow charge delocalization via resonance increase stability.

110. Which of the following species does not exist?

A. Be₂ B. B₂ C. C₂ D. N₂
 A. $(\sigma_{1s})^2(\sigma_{1s}^*)^2(\sigma_{2s})^2(\sigma_{2s}^*)^2$; Does not exist.

B. $(\sigma_{1s})^2(\sigma_{1s}^*)^2(\sigma_{2s})^2(\sigma_{2s}^*)^2 \left\{ \begin{matrix} (\pi_{2p_y})^1 \\ (\pi_{2p_z})^1 \end{matrix} \right\}$; Exist.

C. $(\sigma_{1s})^2(\sigma_{1s}^*)^2(\sigma_{2s})^2(\sigma_{2s}^*)^2 \left\{ \begin{matrix} (\pi_{2p_y})^2 \\ (\pi_{2p_z})^2 \end{matrix} \right\}$; Exist.

D. $(\sigma_{1s})^2(\sigma_{1s}^*)^2(\sigma_{2s})^2(\sigma_{2s}^*)^2(\sigma_{2p_x})^2 \left\{ \begin{matrix} (\pi_{2p_y})^2 \\ (\pi_{2p_z})^2 \end{matrix} \right\}$; Exist



111. NF₃, BF₃ and ClF₃ all have molecular formula of the type XF₃ which molecules has Trigonal pyramidal molecules geometry?

A. NF₃ and BF₃ B. NF₃ and ClF₃ C. BF₃ and ClF₃ D. NF₃

Brief:

❖ **BF₃ (Boron Trifluoride)**

- ✓ Boron has 3 valence electrons and forms three bonds with fluorine.
- ✓ There are no lone pairs on boron.
- ✓ The molecule adopts a Trigonal Planar (120°) geometry.

❖ **NF₃ (Nitrogen Trifluoride)**

- ✓ Nitrogen has 5 valence electrons and forms three bonds with fluorine.
- ✓ One lone pair remains on nitrogen.
- ✓ This results in a Trigonal Pyramidal geometry (similar to ammonia, NH₃).
- ✓ The bond angles are slightly less than 109.5° due to lone pair repulsion.

❖ **ClF₃ (Chlorine Trifluoride)**

- ✓ Chlorine has 7 valence electrons and forms three bonds with fluorine.
- ✓ Two lone pairs remain on chlorine.
- ✓ This results in a T-shaped geometry, not trigonal pyramidal.

112. According to the molecular orbital theory, which one of the following is not paramagnetic?

A. F₂ B. H₂ C. NO⁺ D. NO

Brief:

According to Molecular Orbital (MO) Theory, a molecule is paramagnetic if it has unpaired electrons in its molecular orbitals. Conversely, it is diamagnetic if all its electrons are paired.

O_2 (Oxygen molecule)

- MO Configuration: $(\sigma_{2s})^2(\sigma_{2s}^*)^2(\sigma_{2p_z})^2(\pi_{2p_x})^2(\pi_{2p_y})^2(\pi_{2p_x}^*)^1(\pi_{2p_y}^*)^1$
- Paramagnetic (due to unpaired electrons in π^* orbitals).

B_2 (Boron molecule)

- MO Configuration: $(\sigma_{2s})^2(\sigma_{2s}^*)^2(\pi_{2p_x})^1(\pi_{2p_y})^1$
- Paramagnetic (due to unpaired electrons in π_{2p_x} and π_{2p_y}).

C_2 (Carbon molecule)

- MO Configuration: $(\sigma_{2s})^2(\sigma_{2s}^*)^2(\pi_{2p_x})^2(\pi_{2p_y})^2(\sigma_{2p_z})^2$
- Diamagnetic (all electrons are paired).

N_2 (Nitrogen molecule)

- MO Configuration: $(\sigma_{2s})^2(\sigma_{2s}^*)^2(\pi_{2p_x})^2(\pi_{2p_y})^2(\sigma_{2p_z})^2$
- Diamagnetic (all electrons are paired).

113. All of the these are characteristic of most ionic compounds in the solid phase, EXCEPT

- | | |
|----------------------------------|-------------------------------------|
| A. High melting point | C. Solubility in water |
| B. High electrical conductivity. | D. Insolubility in organic solvents |

Brief:

✚ Properties of ionic compound

Ionic compounds have high melting points, hard and brittle.

Ionic compounds dissociate into ions when dissolved in water.

Solutions and melted ionic compounds conduct electricity, but solid materials do not.

114. What is the bond order of O_2^+ A. 2 B. 2.5 C. 1.5 D. 3

Brief:

The bond order formula is:

$$\text{Bond Order} = \frac{\text{Number of bonding electrons} - \text{Number of antibonding electrons}}{2}$$

- Bonding electrons: 8 (from $\sigma(2p_z)$, $\pi(2p_x)$, and $\pi(2p_y)$)
- Antibonding electrons: 3 (from $\pi^*(2p_x)$ and $\pi^*(2p_y)$)

$$\text{Bond Order} = \frac{8 - 3}{2} = \frac{5}{2} = 2.5$$

115. Which properties are not possessed by metals?

- | | | | |
|-----------------|--------------|---------------------|-------------------------|
| A. Malleability | B. Ductility | C. Hard and Brittle | D. Thermal conductivity |
|-----------------|--------------|---------------------|-------------------------|

Brief:

✚ Physical properties of metals are:

- Metals are malleable and ductile.
- Metals are good conductors of heat and electricity.

- Metals are lustrous (shiny) and can be polished.
- Metals are solids at room temperature (except mercury, which is liquid).
- Metals are tough and strong.

116. How does Molecular Orbital Theory (MOT) explain bonding in molecules that are inconsistent with the predictions of the Valence Bond Theory (VBT)?

- A. MOT suggests that bonding occurs through localized electron pairs between atoms.
- B. MOT explains bonding through delocalized molecular orbitals, which are extended over the entire molecule.
- C. MOT only applies to diatomic molecules and is not useful for larger molecules.
- D. MOT assumes all electrons are localized in atomic orbitals, similar to VBT.

Brief:

- * MOT explains bonding through delocalized molecular orbitals, which are extended over the entire molecule.

117. In what way does the application of VSEPR Theory provide a more accurate geometric structure prediction than Lewis Theory for molecules with multiple resonance forms

- A. VSEPR Theory predicts bond angles by considering only the number of atoms in the molecule.
- B. VSEPR Theory accounts for electron pair repulsion and can predict the actual geometry in molecules with resonance, while Lewis Theory relies on static bond representations.
- C. VSEPR Theory only predicts the linear geometry of molecules.
- D. VSEPR Theory assumes that all resonance forms exist simultaneously in a molecule.

Brief:

VSEPR Theory accounts for electron pair repulsion and can predict the actual geometry in molecules with resonance, while Lewis Theory relies on static bond representations.

118. How does Valence Bond Theory (VBT) explain the concept of hybridization, and how is this concept extended in Molecular Orbital Theory (MOT)?

- A. VBT uses hybridization to form molecular orbitals, while MOT does not consider hybridization.
- B. VBT explains bonding through the overlap of hybridized atomic orbitals to form localized bonds, while MOT involves delocalized molecular orbitals that span the entire molecule.
- C. VBT focuses on molecular orbitals, while MOT emphasizes hybridization of atomic orbitals.
- D. VBT and MOT are identical in their approach to hybridization.

Brief:

- * VBT explains bonding through the overlap of hybridized atomic orbitals to form localized bonds, while MOT involves delocalized molecular orbitals that span the entire molecule.

119. What is the role of bond order in Molecular Orbital Theory (MOT) and how does it relate to the bond predictions made by Lewis structures and VSEPR Theory?

- A. Bond order is irrelevant in MOT and does not affect bonding predictions.
- B. Bond order in MOT is calculated as $(\text{bonding electrons} - \text{antibonding electrons}) / 2$, providing a quantitative measure of bond strength, which is not directly addressed by Lewis structures or VSEPR Theory.
- C. Bond order is only relevant in Lewis structures and VSEPR Theory.
- D. VSEPR Theory provides bond order predictions through electron pair geometry.

Brief:

- * Bond order in MOT is calculated as $(\text{bonding electrons} - \text{antibonding electrons}) / 2$, providing a quantitative measure of bond strength, which is not directly addressed by Lewis structures or VSEPR Theory

120. How do the limitations of the Valence Bond Theory (VBT) in explaining the magnetic properties of molecules contrast with the insights provided by Molecular Orbital Theory (MOT)?

- A. VBT predicts paramagnetism in molecules with unpaired electrons, while MOT predicts diamagnetism.
- B. VBT cannot explain the magnetic properties of molecules, while MOT accounts for unpaired electrons in antibonding molecular orbitals, explaining paramagnetism.
- C. VBT always correctly predicts magnetic properties of molecules.
- D. MOT does not address magnetic properties at all.

Brief:

- * VBT cannot explain the magnetic properties of molecules, while MOT accounts for unpaired electrons in antibonding molecular orbitals, explaining paramagnetic.

2.3. Self-practice questions

121. Which of the following factors contributes most significantly to the strength of a covalent bond
- A. Atomic size
 - B. Bond order
 - C. Electronegativity difference
 - D. Ionization energy
122. Which of the following correctly describes the relationship between bond order and bond length?
- A. As bond order increases, bond length increases
 - B. As bond order increases, bond length decreases
 - C. Bond order has no effect on bond length
 - D. As bond order decreases, bond length decreases
123. In a diatomic molecule, the bond is considered purely covalent when the electronegativity difference is:
- A. 0
 - B. Between 0.5 and 1.5
 - C. Between 1.5 and 2.0
 - D. Greater than 2.0
124. Which of the following compounds exhibits the most ionic character?
- A. NaCl
 - B. H₂O
 - C. CO₂
 - D. NH₃
125. What is the correct order of bond energy for the following bonds?
- A. C-C < C=C < C≡C
 - B. C=C < C-C < C≡C
 - C. C≡C < C=C < C-C
 - D. C=C < C≡C < C-C
126. Which type of bonding is primarily responsible for the high electrical conductivity in metals
- A. Covalent bonding
 - B. Ionic bonding
 - C. Metallic bonding
 - D. van der Waals forces
127. Which of the following statements about polar covalent bonds is true?
- A. They involve complete transfer of electrons
 - B. They occur between atoms of equal electronegativity
 - C. They involve partial sharing of electrons with an uneven charge distribution
 - D. They do not contribute to molecular polarity
128. Which of the following bonds would be the strongest?
- A. N-N
 - B. N=N
 - C. N≡N
 - D. N-O
129. Which of the following is NOT a characteristic of covalent compounds?
- A. Low melting and boiling points
 - B. Poor electrical conductivity in the solid state
 - C. High solubility in polar solvent
 - D. Formation of discrete molecules
130. Which of the following elements would form the most stable diatomic molecule?
- A. Oxygen
 - B. Fluorine
 - C. Nitrogen
 - D. Hydrogen
131. The unusually high boiling point of water compared to other hydrides is primarily due to:
- A. van der Waals forces
 - B. Dipole-dipole interactions
 - C. Hydrogen bonding
 - D. London dispersion forces
132. Which of the following compounds exhibits the strongest hydrogen bonding?
- A. CH₄
 - B. NH₃
 - C. H₂S
 - D. PH₃
133. Which of the following factors increases the strength of London dispersion forces?
- A. Increased molecular polarity
 - B. Increased molecular weight
 - C. Decreased molecular volume
 - D. Increased bond order

134. Which intermolecular force is responsible for the solubility of nonpolar gases in water?
- Dipole-dipole interactions
 - London dispersion forces
 - Hydrogen bonding
 - Ion-dipole forces
135. Which of the following molecules exhibits both dipole-dipole forces and hydrogen bonding?
- CH₃Cl
 - CH₄
 - HF
 - CO₂
136. Which of the following molecules is expected to have the highest boiling point?
- H₂S
 - H₂O
 - CO₂
 - CH₄
137. Which type of intermolecular force explains why noble gases can be liquefied at low temperatures?
- Dipole-dipole interactions
 - London dispersion forces
 - Hydrogen bonding
 - Ion-dipole interactions
138. Which factor most strongly influences the lattice energy of an ionic compound?
- Bond angle
 - Ion size and charge
 - Atomic number
 - Electronegativity difference
139. Which of the following compounds has the highest lattice energy?
- NaCl
 - MgO
 - KBr
 - CaF₂
140. Ionic compounds tend to be soluble in water because:
- They are highly conductive in solid state
 - Water is a nonpolar solvent
 - The strong ion-dipole forces overcome the lattice energy
 - Water molecules form covalent bonds with ionic compounds
141. Which property is NOT a characteristic of ionic compounds?
- High melting points
 - Electrical conductivity in solid state
 - High solubility in polar solvents
 - Formation of crystalline structures
142. Which of the following has the most covalent character?
- NaCl
 - AlCl₃
 - KBr
 - LiF
143. The formation of an ionic bond is typically accompanied by:
- Energy absorption
 - Energy release
 - No energy change
 - Bond length increase
144. Which molecule has resonance structures?
- H₂O
 - CH₄
 - O₃
 - NH₃
145. The formal charge on the central atom in SO₄²⁻ is:
- 0
 - +1
 - 1
 - +2
146. Which of the following species does NOT follow the octet rule?
- BF₃
 - NH₃
 - CH₄
 - H₂O
147. Which statement about resonance is correct?
- Resonance structures exist in equilibrium
 - Resonance is a rapid exchange of electrons between bonds
 - Resonance structures depict different possible electron arrangements
 - Resonance results in fractional bond orders
148. The correct Lewis structure for NO₃⁻ contains how many resonance structures?
- 1
 - 2
 - 3
 - 4
149. Which molecule has a formal charge of zero on all atoms?
- HCN
 - CO₂
 - NO₂
 - O₃
150. Which type of hybridization occurs in the central atom of a molecule with a tetrahedral geometry?
- sp
 - sp²
 - sp³
 - sp³d
151. Which of the following molecules exhibits sp² hybridization on the central atom?
- CO₂
 - CH₄
 - BF₃
 - NH₃
152. The hybridization of the central atom in XeF₄ is:
- sp²
 - sp³
 - sp³d
 - sp³d²

153. Which of the following molecules has a bent molecular geometry due to lone pairs?
 A. CO_2 B. H_2O C. BeCl_2 D. BF_3
154. Which of the following has a bond angle of approximately 120° ?
 A. H_2O B. CH_4 C. BF_3 D. NH_3
155. The shape of ClF_3 according to the VSEPR theory is:
 A. Trigonal planar C. T-shaped
 B. Trigonal pyramidal D. Square pyramidal
156. Which of the following species has a square planar geometry?
 A. BF_3 B. XeF_4 C. PCl_5 D. CH_4
157. A molecule with sp^3d^2 hybridization and no lone pairs on the central atom has which molecular shape?
 A. Octahedral C. See-saw
 B. Trigonal bipyramidal D. T-shaped
158. Which of the following species has an expanded octet?
 A. CO_2 B. NH_3 C. SF_6 D. CH_4
159. Which of the following statements about VSEPR theory is incorrect?
 A. Lone pairs repel bonding pairs more strongly than bonding pairs repel each other
 B. Multiple bonds have greater repulsive force than single bonds
 C. Bond angles in NH_3 are exactly 120°
 D. Electron domains determine molecular shape
160. Which of the following species has an unpaired electron in its molecular orbital diagram?
 A. N_2 B. O_2 C. F_2 D. CO
161. According to molecular orbital theory, which of the following statements is correct for O_2 ?
 A. It is diamagnetic
 B. It has a bond order of 3
 C. It has unpaired electrons
 D. The π orbitals are filled before the σ orbitals
162. Which of the following species has a bond order of 2.5?
 A. NO B. O_2 C. N_2^+ D. CO
163. Which of the following correctly represents the molecular orbital filling order?
 A. $\sigma(2s) < \sigma^*(2s) < \pi(2p) < \sigma(2p) < \pi^*(2p) < \sigma^*(2p)$
 B. $\sigma(2s) < \pi(2p) < \sigma^*(2s) < \sigma(2p) < \pi^*(2p) < \sigma^*(2p)$
 C. $\pi(2p) < \sigma(2s) < \sigma^*(2s) < \sigma(2p) < \pi^*(2p) < \sigma^*(2p)$
 D. $\sigma(2s) < \sigma^*(2s) < \sigma(2p) < \pi(2p) < \pi^*(2p) < \sigma^*(2p)$
164. What is the total number of molecular orbitals formed when two oxygen atoms combine?
 A. 6 B. 10 C. 12 D. 16
165. Which of the following molecules has delocalized π electrons?
 A. CH_4 B. H_2O C. CO_2 D. Benzene
166. Which of the following species has the highest bond order?
 A. O_2 B. N_2 C. NO D. CO
167. Which molecular orbital is the highest energy occupied molecular orbital (HOMO) in a typical diatomic oxygen molecule (O_2)?
 A. $\sigma(2s)$ B. $\pi(2p)$ C. $\pi^*(2p)$ D. $\sigma^*(2p)$
168. According to valence bond theory, the sidewise overlap of p orbitals leads to the formation of
 A. Sigma (σ) Bonds B. Pi (π) Bonds C. Coordinate bonds D. Ionic bonds
169. Which of the following molecules is paramagnetic?
 A. CO B. O_2 C. N_2 D. F_2
170. How many number of pi and sigma electrons of toluene are ____ and ____ respectively
 A. 10 and 12 B. 6 and 32 C. 6 and 15 D. 3 and 10

UNIT THREE

3. CHEMICAL REACTION AND STOICHIOMETRY

Unit Outcomes

- *At the end of this unit, you will be able to*
- ❖ *Explaining different types of reaction: combination reaction, decomposition reaction, single displacement reaction, double displacement reaction, combustion reaction*
 - ❖ *Develop skills in writing and balancing chemical equations;*
 - ❖ *Identifying the reaction whether it is redox or non-redox reaction*
 - ❖ *Developing skills of determining oxidation number for redox reaction*
 - ❖ *Analyses redox reactions by specifying the oxidizing agent, the reducing agent, the substance reduced or oxidized*
 - ❖ *Analyzing theoretical yield, actual yield and percentage yield*
 - ❖ *Develop skills in solving problems based on chemical equations (mole-mole, mass - mass, volume - volume and mass - volume problems);*

3.1. Introduction

3.1.1 Chemical Reactions and Stoichiometry

Chemical Reactions are processes in which one or more substances are converted into new substances with different chemical properties. Stoichiometry is the study of the quantitative relationships between the reactants and products in a chemical reaction.

3.1.2. Types of Chemical Reactions

Type of Reaction	Description	Example
Combination	Two or more reactants combine to form a single product.	$A+B\rightarrow AB$
Decomposition	A single reactant breaks down into two or more products.	$AB\rightarrow A+B$
Single Displacement	One element displaces another in a compound.	$A+BC\rightarrow AC+B$
Double Displacement	Two compounds exchange ions to form new compounds.	$AB+CD\rightarrow AD+CB$
Combustion	A substance reacts with oxygen to produce heat and light.	$C_xH_y+O_2\rightarrow CO_2+H_2O$
Redox (Oxidation-Reduction)	A reaction involving the transfer of electrons.	$2Na+Cl_2\rightarrow 2NaCl$

3.1.3. Oxidation-Reduction (Redox) Reactions

In redox reactions, electrons are transferred between reactants. One substance undergoes **oxidation** (loses electrons), and the other undergoes **reduction** (gains electrons).

- **Oxidation:** Increase in oxidation state (loss of electrons).
- **Reduction:** Decrease in oxidation state (gain of electrons).

3.1.4. Fundamental Laws of Chemical Reactions

Law	Description
Law of Conservation of Mass	Mass is neither created nor destroyed in a chemical reaction.
Law of Definite Proportions	A chemical compound always contains the same elements in the same proportions by mass.
Law of Multiple Proportions	When two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other are in simple whole number ratios.

3.1.5. Molecular Formula vs. Empirical Formula

- Molecular Formula: Shows the exact number of atoms of each element in a molecule.
 - ✓ Example: Glucose $C_6H_{12}O_6$
- Empirical Formula: The simplest whole-number ratio of elements in a compound.
 - ✓ Example: The empirical formula of glucose is CH_2O

To determine the empirical formula from the molecular formula:

1. Find the greatest common factor (GCF) of the subscripts in the molecular formula.
2. Divide each subscript by the GCF.

3.1.7. Endothermic and Exothermic Reaction:

Exothermic Reaction:

- Definition: An endothermic reaction absorbs energy (heat) from its surroundings.
- Energy change: The products of the reaction have more energy than the reactants.
- Temperature change: The surrounding temperature decreases because the reaction takes in heat.
- Example: Photosynthesis in plants, where energy from sunlight is absorbed to convert carbon dioxide and water into glucose and oxygen. General formula: $\text{Reactants} + \text{Energy} \rightarrow \text{Products}$

Exothermic Reaction:

- Definition: An exothermic reaction releases energy (heat) into the surroundings.
- Energy change: The reactants have more energy than the products.
- Temperature change: The surrounding temperature increases because the reaction releases heat.
- Example: Combustion of fuel, like burning wood or gasoline, which releases energy in the form of heat and light. General formula: $\text{Reactants} \rightarrow \text{Products} + \text{Energy}$
 - Endothermic: Energy is absorbed, making it feel cooler
 - Exothermic: Energy is released, making it feel warmer.

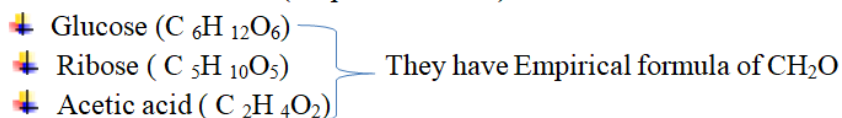
3.2. Questions with their explanation

171. Which of group of compounds are similar in their empirical formula but different molecular formula?

- A. Glucose ($C_6H_{12}O_6$), ribose ($C_5H_{10}O_5$), Acetic acid ($C_2H_4O_2$)
- B. Benzene (C_6H_6), formaldehyde (CH_2O), ribose ($C_5H_{10}O_5$)
- C. Glucose ($C_6H_{12}O_6$), ribose ($C_5H_{10}O_5$), Maltose ($C_{12}H_{22}O_{11}$)
- D. Ribose ($C_5H_{10}O_5$), Maltose ($C_{12}H_{22}O_{11}$), Glycogen ($C_{24}H_{42}O_{21}$)

Brief:

Molecular formula = n (Empirical formula)



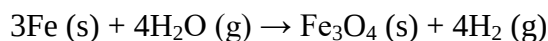
172. Which one of the following properties of a substance does NOT represent an intensive physical property?

- A. Melting point
- B. Hardness
- C. Density
- D. Volume

Brief:

Property	Definition	Examples	Characteristics
Intensive	A property that does not depend on the amount of substance present.	- Temperature - Pressure - Density - Refractive Index - Color - Melting Point	- Independent of the size or amount of system. - Used to identify substances. - Can be uniform or vary with the system.
Extensive	A property that depends on the amount of substance of the system.	- Mass - Volume - Energy - Enthalpy - Entropy - Length	- Depends on the amount or size of the material. - Cannot be used to identify substances. - Additive for different parts of the system.

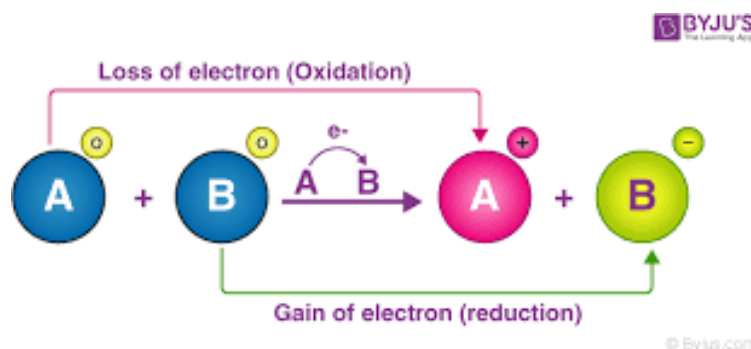
173. Which of the following statements about the given reaction are correct?



- (i) Iron metal is getting oxidized (ii) Water is getting reduced
(iii) Water is acting as reducing agent (iv) Water is acting as oxidising agent
- A. (i), (ii) and (iii) C. (iii) and (iv)
B. (i), (ii) and (iv) D. (ii) and (iv)

Brief:

- ✓ Oxidation-reduction reaction is a type of chemical reaction that involves a transfer of electrons between two species.
- ✓ It is any chemical reaction in which the oxidation number of a molecule, atom, or ion changes by gaining or losing an electron.
 - Oxidation reaction:
 - ✓ It is half reaction that to be increase its oxidation number
 - ✓ It is the cause of other substance to be reduced, so called reducing agent.
 - Reduction reaction
 - ✓ It is half reaction that to be decrease its oxidation number.
 - ✓ It is the cause of other substance to be oxidized, so called oxidizing agent.



174. The type of the reactions in which two reacting substances are equally consumed to give product is

- A. Limiting reactant C. Excess reactant
B. Stoichiometric proportion D. Theoretical yield

Brief:

- ✚ During the course of chemical reaction, the reacting species called reactants may or may not consume equally. When the consumption of reactant and product are not yet equally consumed, one is consumed first (called limiting reactant) and other to be left excess (excess reactant).

✚ When the ratio of moles of reactants entering a reactor is in the same ratio as in a balanced chemical reaction, (the reacting species are consuming equally) then the reactants are said to be in stoichiometric proportion.

175. In an experiment 1.056 g of a metal carbonate, containing an unknown metal M, is heated to give the metal oxide and 0.376 g of CO₂. $\text{MCO}_3 (\text{s}) + \text{heat} \rightarrow \text{MO} (\text{s}) + \text{CO}_2 (\text{g})$. Then the metal and its molar mass produced are ____ and ____ respectively.

- A. Copper and 63.5g/mol C. Zinc and 60g/mol
B. Lead and 78g/mol D. Chromium and 56 g/mol

Brief

- ✓ Mass of metal carbonate = 1.056 g
- ✓ Mass of CO₂ = 0.376 g
- ✓ We are required to identify the metal M in the carbonate
- ✓ The equation for the combustion of the carbonate;
- ✓ $\text{MCO}_3 (\text{s}) + \text{heat} \rightarrow \text{MO} (\text{s}) + \text{CO}_2 (\text{g})$

➤ Calculating the number of moles of CO₂

- ✓ Moles = Mass ÷ molar mass
- ✓ Molar mass CO₂ = 44.01 g/mol
- ✓ Moles = 0.376 g ÷ 44.01 g/mol
- = **0.00854 Moles**
- Determining the moles of MCO₃
- ✓ 1 mole of the metal carbonate decomposes to yield 1 mole CO₂
- ✓ Therefore; Mole ratio of MCO₃ : CO₂ = 1 : 1
- ✓ Hence; Moles of MCO₃ = **0.008 mole**

➤ Determining the molar mass of MCO₃

- ✓ Moles of MCO₃ = 0.00854 moles
- ✓ Mass of the carbonate = 1.056 g
- ✓ Molar mass = Mass ÷ moles
= 1.056 g ÷ 0.00854 moles
= **123.65 g/mol**

➤ Atomic mass of the metal

- ✓ The molar mass of metal carbonate = 123.65 g/mol
- ✓ Atomic mass of carbon = 16
- ✓ Atomic mass of oxygen = 12.01
- ✓ Therefore;
123.65 g = M + 12.01 + (16 × 3)
M = 63.5 g
- ✓ The atomic mass of M is 63.64 approximately the same of copper.
- ✓ Atomic mass of Cu is 63.546
- ✓ Therefore, identity of **M is Cu**

176. A compound is found to be 40.0% carbon, 6.7% hydrogen and 53.5% oxygen.

Its molecular mass is 60. g/mol. What is its molecular formula?

- A. C₂H₄O₂ B. C₂H₄O₂ C. C₂H₄O₂ D. C₂H₄O₂

The empirical formula is CH_2O , and the molecular formula is some multiple of this.

Explanation:

In 100 g of the unknown, there are $\frac{40.0 \cdot g}{12.011 \cdot g \cdot mol^{-1}}$ C; $\frac{6.7 \cdot g}{1.00794 \cdot g \cdot mol^{-1}}$ H;
and $\frac{53.5 \cdot g}{16.00 \cdot g \cdot mol^{-1}}$ O.

We divide thru to get, $C:H:O = 3.33:6.65:3.34$. When we divide each elemental ratio by the LOWEST number, we get an empirical formula of CH_2O , i.e. near enough to WHOLE numbers.

Now the molecular formula is always a multiple of the empirical formula; i.e. $(EF)_n = MF$.

So $60.0 \cdot g \cdot mol^{-1} = n \times (12.011 + 2 \times 1.00794 + 16.00)g \cdot mol^{-1}$.

177. **From the following statements choose one which is wrong?**

- A. Theoretical yield is obtained from the laboratory work
- B. Actual yield is the one which is on hand
- C. Theoretical yield is greater than the actual yield
- D. Theoretical yield is obtained by calculation from the balanced chemical equation.

Brief:

- ❖ The theoretical yield is the amount expected based on the amount of reactants added and the balanced chemical equation.
- ❖ The actual yield is the amount of product that was produced with losses and less than ideal conditions in place.

Comparison	Actual yield	Theoretical yield
Definition	The amount of product obtained after performing a chemical reaction	The amount of product calculated from a balance chemical equation
Also known as (A.K.A)	Experimental yield	Expected yield
Value	Less than theoretical yield	More than actual yield
Basis	Experimental	Theoretical
Formula	$A.Y = P.Y / T.Y \times 100\%$	$T.Y = A.Y / P.Y \times 100\%$

178. Which of the following reactions, one is not a redox reaction?

- A. $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ C. $\text{NaCl} + \text{KBr} \rightarrow \text{KCl} + \text{NaBr}$
 B. $\text{Mg} + \text{Cl}_2 \rightarrow \text{MgCl}_2$ D. $\text{Mg} + \text{CuCl}_2 \rightarrow \text{MgCl}_2 + \text{Cu}$

Brief:

- ✓ Redox and nonredox reactions are two major types of chemical reactions. These two groups differ from each other depending on the changes that happen to the oxidation states of the chemical elements from which the reactants are made of.
- ✓ When a change in oxidation number occurs in a reaction, with both an increase in oxidation number (oxidation) and a decrease in oxidation number (reduction), then the reaction is classified as redox.
- ✓ If the reaction takes place without transfer of electron (without change in oxidation number) occurs, then the reaction is non-redox.
- ✓ E.g Metathesis reaction, Acid-base reaction.

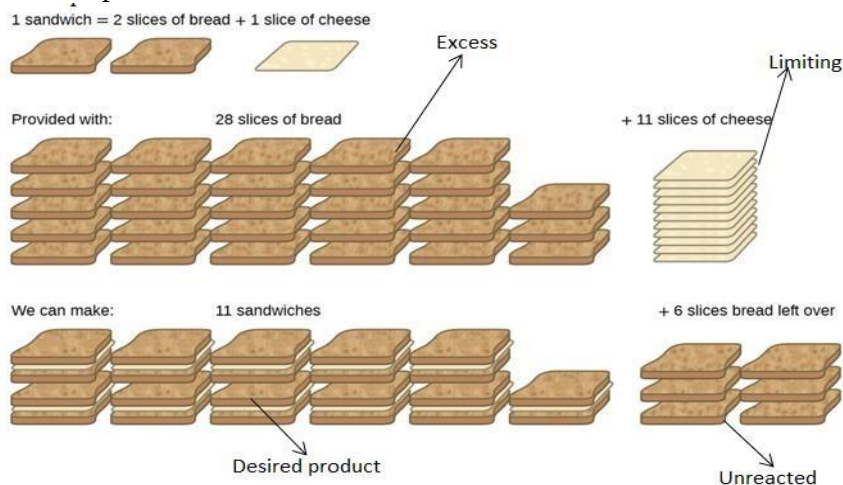
179. Which one of the following is not true about limiting and excess reactants?

- A. The reactant that gives smaller amount of the product is limiting reactants
 B. The reactants left unreacted at the end of the reaction is excess reactant
 C. The reactant that determines the amount of product formed is excess reactants
 D. The reactant that determines the amount of the product formed is limiting reactant

Brief:

- ✓ The limiting reagent in a chemical reaction is the reactant that will be consumed completely.
- ✓ Once there is no more of that reactant, the reaction cannot proceed.
- ✓ Therefore it limits the reaction from continuing, said to be **reaction determinant**
- ✓ The excess reagent is the reactant that could keep reacting if the other had not been consumed.

Example



180. In a polyionic compound, the sum of the oxidation numbers of all the elements equals
A. Zero B. +1 C. -1 D. The charge of the substance

Brief:

- ✓ The sum of the oxidation numbers in a polyatomic ion is equal to the net charge of that respective ion.
- ✓ For polyatomic cation, the sum of oxidation number of all atoms is equals to the positive charge of ion.
- ✓ For polyatomic anion, the sum of oxidation number of all atoms is equals to the negative charge of ion.
- ✓ E.g. If a polyatomic ions has a -1 net charge, then the sum of its elements' oxidation numbers must be -1; likewise, if a polyatomic ions has a +2 net charge, then the sum must be +2.

181. What mass of aluminum would be completely oxidized by 44.8 L of oxygen to produce aluminum oxide at STP? (Al=27g/mol, O=16g/mol)

- A. 50.4g B. 72g C. 59.73g D. 216g

Brief:

- ✓ Aluminium will undergo oxidation reaction in presence of oxygen to give Aluminium oxide
 - ✓ Write balanced chemical equation
 - ✓ $4Al + 3O_2 \rightarrow 2Al_2O_3$
 - ✓ Find the mole of Al using mole-volume relation ship
- $$\frac{\text{Mole of Al}}{\text{mol of Al (4 mole)}} = \frac{\text{volume of } O_2 (44.8)}{\text{volume of } O_2 (3 \times 22.4L)}$$

- ✓ Convert mole of Al to mass,
- ✓ $\text{Mass} = \text{Mole of Al} \times \text{molar mass of Al}.$

106. If 20 g of CaCO_3 and 25 g of HCl are mixed together, what mass of the CO_2 is produced according to the following reaction? ($\text{Ca}=40\text{g/mol}$, $\text{C}=12\text{g/mol}$, $\text{O}=16\text{g/mol}$)

- A. 15.07g B. 14.6g C. 8.8g D. 10.4g

Brief:

- Convert the word equation in to chemical equation
- Balance chemical equation: $\text{CaO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$

Let CaO_3 be completely consumed. 20g <u>m</u> $\text{CaO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$ 100g/mol 44 g/mol $\frac{20 \text{ g}}{100 \text{ g/mol}} = \frac{m}{44 \text{ g/mol}}$ $m = 20 \text{ g} \times 44 \text{ g/mol} / 100 \text{ g/mol} = \mathbf{8.8 \text{ gram}}$	Let HCl be completely consumed. 25 g m $\text{CaO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$ 72g/mol 44g/mol $\frac{25 \text{ g}}{72 \text{ g/mol}} = \frac{m}{44 \text{ g/mol}}$ $m = 25 \text{ g} \times 44 \text{ g/mol} / 72 \text{ g/mol} = \mathbf{15.3 \text{ gram}}$
--	---

The reactant that provides small amount of product is limiting reactant.

So, CaO_3 is limiting reactant while HCl is Excess.

$CaCO_3$ is reaction determinant.

182. Which of the following statements is true about reducing and oxidizing agent?

- A. The substance that oxidizes other substances is reducing agent
- B. The compound that contains the oxidized substance is oxidizing agent
- C. The element that lost its electron is reducing agent
- D. The element that gains electron is said to be reducing agent

Brief:

- A **reducing agent** (or **reductant**) is a substance that **loses electrons** in a redox reaction. By losing electrons, it gets **oxidized** itself and helps another substance gain electrons (get reduced).
- An **oxidizing agent** (or **oxidant**) is a substance that **gains electrons**, meaning it gets **reduced** itself while causing another substance to lose electrons (get oxidized).

183. For the reaction $3\text{Cu} + 8\text{HNO}_3 \rightarrow 3\text{Cu}(\text{NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}$, what is reduced substance and oxidizing agent respectively?

- A. Copper, Nitrogen
- B. Nitrogen, Copper
- C. Nitrogen, HNO_3
- D. Copper, HNO_3

Brief:

➤ Identify Oxidation States

- ✓ Copper (Cu) in its elemental form has an oxidation state of 0.
- ✓ In $\text{Cu}(\text{NO}_3)_2$, copper has an oxidation state of +2.
- ✓ Nitrogen in HNO_3 (in the NO_3^- ion) has an oxidation state of +5.
- ✓ Nitrogen in NO has an oxidation state of +2.

➤ Identify Oxidation and Reduction

- ✓ Copper (Cu) goes from 0 to +2 → It loses electrons (oxidation).
- ✓ Nitrogen in HNO_3 goes from +5 to +2 → It gains electrons (reduction).

➤ Identify the Reduced Substance and Oxidizing Agent

- ✓ The substance that is reduced is Nitrogen (from HNO_3 , changing from +5 to +2).
- ✓ The oxidizing agent is the substance that causes oxidation, which is HNO_3 because it contains the nitrate ion (NO_3^-), which accepts electrons.

184. Of the following statements which one is wrong?

- A. Chemical reaction comes before chemical equation
- B. Less active metal displaces more active metal
- C. All chemical reactions are chemical changes
- D. All decomposition reactions are endothermic reactions

Brief:

- A chemical reaction occurs first, and then we represent it using a chemical equation.
- A less active metal cannot displace a more active metal from its compound. A more reactive metal replaces a less reactive metal in a displacement reaction.
- Every chemical reaction leads to a chemical change, forming new substances with different properties.
- While many decomposition reactions require heat (endothermic), some, like the electrolysis of water and some explosive decomposition, are exothermic.

185. The type of reaction in which two compounds exchange their ions to form two new compounds is

- | | |
|---------------------------|---------------------------------|
| A. Decomposition reaction | C. Double displacement reaction |
| B. Combination reaction | D. Single displacement |

Brief:

- In a double displacement reaction, two compounds exchange their ions to form two new compounds.
- This typically occurs in aqueous solutions and often results in the formation of a precipitate, gas, or water.
- $AB + CD \rightarrow AD + CB$ = $NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$

186. Suppose that 3 moles of carbon and 8 moles of oxygen gases are supplied in the reaction used to produce carbon dioxide. How gram of carbon dioxide is produced?



- | | | | |
|-------|-------|-------|-------|
| A. 11 | B. 22 | C. 44 | D. 16 |
|-------|-------|-------|-------|

Brief:

- Write the Balanced Chemical Equation
 $C + O_2 \rightarrow CO_2$
- ✓ This equation tells us: 1 mole of carbon (C) reacts with 1 mole of oxygen gas (O₂) to form 1 mole of carbon dioxide (CO₂).
- Identify Given Amounts
 - ✓ 3 moles of Carbon (C)
 - ✓ 8 moles of Oxygen gas (O₂)
 - ✓ We need to determine how much CO₂ is produced.
- Determine the Limiting Reactant
 - ✓ The balanced equation tells us that, 1 mole of C reacts with 1 mole of O₂
 - ✓ compare the ratio of reactants available:
 - ✓ We have 3 moles of C, which would require 3 moles of O₂ to fully react.
 - ✓ We have 8 moles of O₂, which is more than enough.
 - ✓ Since carbon is completely used up first, carbon (C) is the limiting reactant.
- Determine the Moles of CO₂ Produced

- ✓ Since 1 mole of C produces 1 mole of CO_2 , and we have 3 moles of C:
- ✓ 3 moles of CO_2 are produced.

➤ Convert Moles of CO_2 to Grams

- ✓ Molar mass of CO_2 = (Mass of CO_2 = moles $\times$ molar mass = $3 \times 44 = 132$ g.

187. Which one of the following set of choices contains the correct increasing order of oxidation state of chlorine from least to greatest?

- A. Cl_2 , HCl , HOCl , HClO_3 C. HCl , Cl_2 , HOCl , HClO_3
 B. HCl , HOCl , Cl_2 , HClO_3 D. Cl_2 , HClO_3 , HOCl , HCl

Brief:

- To determine the correct increasing order of the oxidation state of chlorine, we calculate the oxidation state of chlorine in each compound:

- ✓ Cl_2 (Chlorine gas) Oxidation state = 0 (elemental form)
- ✓ HCl (Hydrochloric acid) Oxidation state of Cl = -1
- ✓ HOCl (Hypochlorous acid) Oxidation state of Cl = +1
(Oxygen is -2, Hydrogen is +1, so Cl must be +1 to balance)
- ✓ HClO_3 (Chloric acid)..... Oxidation state of Cl = +5
(Each oxygen is -2, total = -6, Hydrogen is +1, so Cl must be +5 to balance)
- ✓ Now, arranging these oxidation states in increasing order



188. What is the oxidation state of Cr in $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$?

- A. +6 B. -3 C. +3 D. -6 E. +7

Brief:

- * Assign oxidation states for known elements:

- NH_4^+ (Ammonium ion): The oxidation state of N in NH_4^+ is +1, and H is +1.
- Oxygen (O) in $\text{Cr}_2\text{O}_7^{2-}$ (Dichromate ion): Oxygen typically has an oxidation state of -2.

- * Let the oxidation state of Cr be x:

- The dichromate ion ($\text{Cr}_2\text{O}_7^{2-}$) has an overall charge of -2.
- Oxygen contributes $7(-2) = -14$.
- There are two Cr atoms, so the total contribution of Cr is $2x$.

- * Set up the equation:

$$- 2x + (-14) = -22x + (-14) = -22x + (-14) = -2$$

189. All of the following are non- redox reaction, EXCEPT

- A. $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
- B. $\text{AgNO}_3 + \text{KCl} \rightarrow \text{AgCl} + \text{KNO}_3$
- C. $\text{MgCO}_3 + \text{Li}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{Li}_2\text{CO}_3$
- D. $\text{HCl} + \text{K}_2\text{Cr}_2\text{O}_7 \rightarrow \text{KCl} + 2\text{CrCl}_3 + 7\text{H}_2\text{O}$

Brief:

A redox reaction involves the transfer of electrons, meaning one substance gets oxidized (loses electrons) and another gets reduced (gains electrons).

- $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$: This is a neutralization reaction, where an acid reacts with a base to form salt and water. There is no change in the oxidation states of the elements involved, so this is not a redox reaction.
- $\text{AgNO}_3 + \text{KCl} \rightarrow \text{AgCl} + \text{KNO}_3$: This is a double displacement reaction (precipitation) where silver chloride (AgCl) precipitates out. Again, there is no change in oxidation states, so this is not a redox reaction.
- $\text{MgCO}_3 + \text{Li}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{Li}_2\text{CO}_3$: This is another double displacement reaction with no change in oxidation states, so it is not a redox reaction.
- $\text{HCl} + \text{K}_2\text{Cr}_2\text{O}_7 \rightarrow \text{KCl} + 2\text{CrCl}_3 + 7\text{H}_2\text{O}$: This is a redox reaction. The chromium (Cr) in $\text{K}_2\text{Cr}_2\text{O}_7$ (potassium dichromate) is reduced from an oxidation state of +6 to +3 in CrCl_3 , and hydrogen is oxidized (from HCl , as it forms water). Therefore, this is the only redox reaction among the options.

3.3. SELF-PRACTICE Questions

190. Which of the following is a chemical change?
A. Melting of ice B. Rusting of iron C. Boiling of water D. Cutting of paper
191. The law of conservation of mass states that:
A. Mass can be created
B. Mass can be destroyed
C. Mass remains constant in a chemical reaction
D. Mass increases after a reaction
192. Which of the following is NOT a correct statement about the limiting reagent?
A. It determines the maximum amount of product formed.
B. It is completely consumed in a chemical reaction.
C. The reagent with the smallest mass is always the limiting reagent.
D. It limits the extent of the reaction.
193. In a combination reaction,
A. A single compound breaks down into two or more products
B. Two or more reactants form a single product
C. Atoms are rearranged without forming new products
D. A substance reacts with oxygen
194. Which of the following is an example of a decomposition reaction?
A. $\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$ C. $\text{Zn} + \text{CuSO}_4 \rightarrow \text{Cu} + \text{ZnSO}_4$
B. $\text{NaCl} \rightarrow \text{Na} + \text{Cl}_2$ D. $\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
195. What type of reaction is represented by the equation:
 $\text{AB} + \text{CD} \rightarrow \text{AD} + \text{CB}$
A. Combination B. Decomposition C. Single replacement D. Double replacement
196. Which of the following is a redox reaction?
A. $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$ C. $\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2$
B. $\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3$ D. $\text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + \text{NaCl}$
197. In a redox reaction, which statement is always true?
A. The oxidizing agent is reduced. C. Oxidation occurs at the cathode.
B. The reducing agent gains electrons. D. Reduction occurs at the anode
198. What is the molar mass of H_2O ?
A. 16 g/mol B. 18 g/mol C. 20 g/mol D. 22 g/mol
199. One mole of any substance contains:
A. 3.01×10^{23} particles
B. 6.02×10^{23} particles
C. 9.11×10^{23} particle
D. 1.67×10^{-24} particles

200. What is the empirical formula of $\text{C}_6\text{H}_{12}\text{O}_6$?
 A. CH_2O B. $\text{C}_3\text{H}_6\text{O}_3$ C. $\text{C}_6\text{H}_{12}\text{O}_6$ D. CHO
201. What volume of hydrogen gas (at STP) is produced when 2 moles of water decompose?

$$2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$$
 A. 11.2 L B. 22.4 L C. 44.8 L D. 67.2 L
202. Which of the following is a limiting reactant?
 A. A reactant that is completely consumed in a reaction
 B. A reactant that remains in excess after the reaction
 C. A product that is formed in excess
 D. A reactant that speeds up the reaction
203. If 5 moles of A react with 7 moles of B according to the equation:

$$2\text{A} + 3\text{B} \rightarrow \text{C}$$
 What is the limiting reagent?
 A. A B. B C. C D. Cannot be determined
204. What is the percentage composition of oxygen in H_2O ?
 A. 50% B. 66.6% C. 88.8% D. 33.3%
205. How many grams of CO_2 are produced when 10 g of CH_4 burns completely?

$$\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$$
 A. 22 g B. 27.5 g C. 30 g D. 40 g
206. What is the oxidation state of sulfur in H_2SO_4 ?
 A. +2 B. +4 C. +6 D. 0
207. How many molecules are present in 36 g of water?
 A. 6.02×10^{23} B. 1.20×10^{24} C. 2.01×10^{23} D. 3.01×10^{23}
208. What mass of NaCl is needed to prepare 500 mL of a 0.5 M solution?
 A. 11.7 g B. 23.4 g C. 58.5 g D. 29.3 g
209. What is the theoretical yield of CO_2 when 10 g of C_3H_8 combusts completely?

$$\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$$
 A. 15 g B. 20 g C. 30 g D. 40 g
210. What is the empirical formula of a compound that contains 40% sulfur, 60% oxygen by mass?
 A. SO B. SO_2 C. S_2O_4 D. SO_3

UNIT FOUR

4. RATE OF REACTION

UNIT OUT COMES

 ***At the end of this unit, you students will be able to***

- ❖ *Explain rate of reaction*
- ❖ *Describe the pre-conditions for a chemical reaction to occur*
- ❖ *List and explain the factors that affect rate of chemical reaction.*
- ❖ *Calculate the rate of appearance and rate of disappearance of some reactions*

4.1. INTRODUCTION

4.1.1. Chemical kinetics

Is the area of chemistry concerned with the speeds, or rates, at which reactions occur? The word “kinetic” suggests movement or change. Chemical kinetics refers to the rate of reaction, which are the changeover times in the concentration of a reactant or a product. Chemical kinetics also the study of rates of chemical processes. This includes investigations of how different experimental conditions can influence the speed of a chemical reaction. The investigations yield information about the reaction's mechanism and transition states. The study of chemical kinetics also uses the construction of mathematical models that can describe the characteristics of a chemical reaction.

Rate of reaction (speed) is not fixed for a particular type of reaction, since it depends on several factors. The rate of reaction gets altered when subjected to small changes in temperature, concentration, pressure etc.

Some chemical reactions proceed quickly, whereas others require days, months or even years to give products. These differences depend primarily on the differences in the chemical nature of the reacting substances. The rate of reaction defined for both reversible and irreversible reactions as well as slow and fast reactions.

Fast reactions

- Combustions of fuel
- Reactions of active metal with acids
- Precipitation reactions
- Neutralization reactions
- Explosions
- Firing a gun
- Petro burning in car engine
- Zn or Mg reacting with an acid
- Acid + Base → Salt + water
- Active metal with water

slow reactions

- Rusting of Iron
- Oxidations coppers
- Souring of milk
- Ripening of fruit
- Rusting
- Fermentation
- Making Cheese
- Formation of Ag₂S
- Oxidation of an apple
- Geological reactions

Rate of reaction is the change in concentration of a reactant or a product over a given period of time.

During the course of a reaction, the concentrations of the reactants decrease while those of the products increase. Thus, the rate of a reaction is calculated by dividing the change in the concentrations of reactants or products by the time taken for the change to occur. For a simple reaction, $A \rightarrow B$: The rate of reaction is expressed in terms of a reactant concentration as

$$r = -\Delta [A]/\Delta t = [A]_f - [A]_i/\Delta t$$

Where r is rate of reaction, A is a reactant, B is a product, $[A]_i$ and $[A]_f$ are initial and final concentration of A respectively, $\Delta [A]$ is change in concentration of A and Δt is the change in time.

- $[A]_f < [A]_i$. Thus, $\Delta[A] = [A]_f - [A]_i$, is negative. This indicates that the reactant is being consumed. However, the rate of a reaction is a positive quantity.

In terms of product concentration, the rate of a reaction is given by:

$$r = \Delta[B] / \Delta t = [B]_f - [B]_i / \Delta t$$

- $[B]_f > [B]_i$. Thus, $\Delta B = [B]_f - [B]_i$ is positive because $[B]$ increases and the rate of a reaction is also positive.
- The unit of rate is $\text{mol L}^{-1} \text{s}^{-1}$ (Ms^{-1}) and the unit of concentration is mole per litre (mol/L)

4.1.2. Reaction Rate and Stoichiometry

Rate expressions in terms of the disappearance of the reactants and the appearance of the products:



A graph of changes in concentration of a reactant and a product over time where time is written as the x-axis and concentration as the y-axis.

In general, for the reaction, $aA + bB \rightarrow cC + dD$

Determination of Rate of Reaction

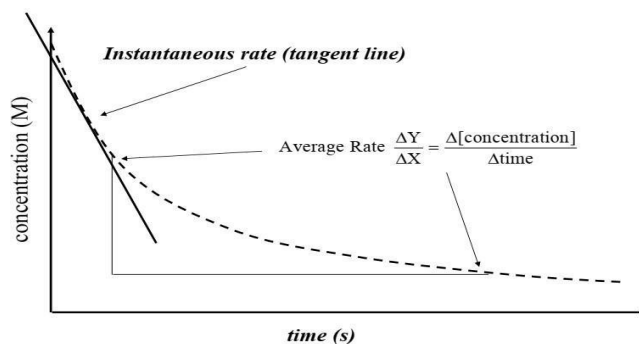
$$r = -\frac{1}{a} \frac{\Delta[A]}{\Delta t} = -\frac{1}{b} \frac{\Delta[B]}{\Delta t} = \frac{1}{c} \frac{\Delta[C]}{\Delta t} = \frac{1}{d} \frac{\Delta[D]}{\Delta t}$$

The rate of reaction can be determined both quantitatively and qualitatively.

Qualitatively: - An idea about the rate of reaction can be obtained by observing either the speed of appearance of the reactants or the speed of appearance (formation) of the /products.

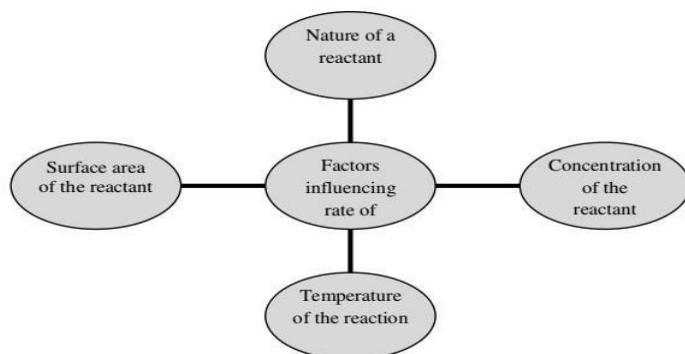
Quantitatively:- Quantity obtained from formula

Instantaneous rate:- is defined as the rate of a reaction for a given instant of time and it is obtained from the slope of the tangent to the concentrations curve at that time



4.1.3. Factors Influencing the Reaction Rates

- In general, different chemical reactions have different rates. Even a chemical reaction involving the same reactants may have different rates under different conditions.



- i. **Temperature of reactants:** - When temperature increases, rate of reaction increases
Reason: - When temperature increases overall kinetic energy of all particles of reactants increases, so more molecules will have energy greater than activation energy. So, the overall frequency of collisions increases, therefore frequency of effective collisions increases, and thus the reaction rate increases
 - ii. **Concentration of reactant:** - The concentration of one or more reactants increases, rate of reaction increases.
Reason: - Increasing the concentration one or more reactants produce more contacts between the reacting particles, which results in increasing the rate of reaction.
 - iii. **For gaseous reactants**, an increase in pressure increases reactions rate
Reason: - Increasing pressure means the volume decreases implies contact (Concentration) increase as result collision increases and reaction will be fast. However, pressure change has no effect on the rate of reaction if the reactants are solids or liquids.
 - iv. **Surface Area of Reactants:** - as surface area increase rate of reaction increases.
Reason:- When the size of the substance subdivisions decreases, surface area increases. It makes the effective collisions increase and therefore the rate of reaction also increases.
- * Why do large crystals of sugar burn more slowly than finely ground sugar?
Answer: - The rate of combustion is proportional to the surface area of sugar exposed to oxygen. Smaller crystals have more surface area and burn faster.
 - * How do homogeneous catalysts and heterogeneous catalyst differ?
A homogeneous catalyst is a catalyst that exists in the same phase (liquid or gas) as the reactants.
 - * A Heterogeneous catalyst, on the other hand, is a catalyst that exists in a different phase from the reactants.

- v. Nature of reactants:-The rate of a reaction depends on the chemical nature of the substances participating in the reaction. A more active a substance the reactive
- V. ***Effect of Catalyst on Rate of Reaction.*** A catalyst is a substance that changes reaction rate by providing a different reaction mechanism one with lower activation energy

Characteristics of catalyst

- Speeds up the reaction forward and backward equally
- May even form temporary chemical bonds with the reactants
- Recovered unchanged in original form at the end of the reaction
- Does not alter the position of equilibrium
- It provides an alternate mechanism by reducing activation energy
- Does not consumed itself in the reactions

4.2. Questions with explanation

211. Consider the hypothetical reaction.



What is the rate of formation of $XY_2(g)$

- A. $-5.0 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$ C. $2.5 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$
B. $5.0 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$ D. $-2.5 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$

Brief:

➤ The formation of the product is half of the rate of the loss of Y. It is correct to change the sign to positive, because we are representing the formation of product, and the negative sign on the reactant to represent loss.

- ✓ Rate of appearance of $XY_2 = \frac{1}{2}$ of disappearance of reactant Y.
- ✓ $-\Delta[Y]/\Delta t = \Delta[XY_2]/\Delta t$
- ✓ $\Delta[XY_2]/\Delta t = -1/2 (5.0 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1})$

212. For the chemical reaction $H_2 + Br_2 \rightarrow 2HBr$, what can be said of the relative rates of consumption of the reactants when compared to the formation of product?

- A. HBr will be produced at half the rate that bromine is consumed.
B. HBr will be produced at the same rate that bromine is consumed.
C. Br_2 and H_2 will be consumed at different rates.
D. HBr will be produced at twice the rate that bromine is consumed

Brief

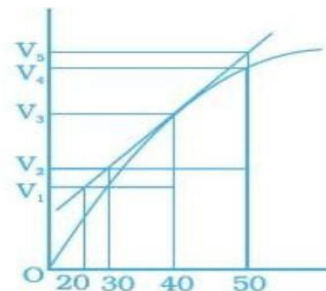
- ✓ The stoichiometry of the overall reaction will give us answers to the questions about comparative rates of formation of products compared to loss of reactants.
- ✓ Rate : $-\Delta[H_2]/\Delta t = -\Delta[Br_2]/\Delta t = 1/2\Delta[HBr]/\Delta t$
- ✓ Bromine and hydrogen are consumed at the same rate while HBr is produced at twice the rate of either Br or H.

213. For the reaction $A + 3B \rightarrow 2C$, how does the rate of disappearance of B compare to the rate of production of C? (self-practice)

- A. The rate of disappearance of B is $1/2$ the rate of appearance of C
B. The rate of disappearance of B is $3/2$ the rate of appearance of C
C. The rate of disappearance of B is $2/3$ the rate of appearance of C
D. The rate of disappearance of B is $1/3$ the rate of appearance of C

214. A graph of volume of hydrogen released v_s time for the reaction between zinc and dil. HCl is given in Fig. On the basis of this mark the correct option is?

- (i) Average rate upto 40s is $\frac{V_3 - V_2}{40}$
- (ii) Average rate upto 40 seconds is $\frac{V_3 - V_2}{40 - 30}$
- (iii) Average rate upto 40 seconds is $\frac{V_3}{40}$
- (iv) Average rate upto 40 seconds is $\frac{V_3 - V_1}{40 - 20}$



Brief:

- Because average rate up to 40 means $V_3 - V_0 / 40 - 0 = V_3 / 40$

215. The rate of a chemical reaction

- Is always dependent of the concentration of all reactants.
- Is always increased with increasing temperatures.
- Is greater with higher activation energies.
- Is independent of surface area.

Brief:

- ✓ The collision model indicates that molecular collision is necessary for reaction. Because an increase in temperature raises molecular velocities and the percentage of effective collisions, the reaction rate increases

216. Which of the following statements is correct?

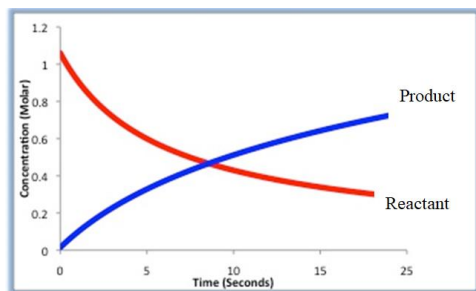
- The rate of a reaction decreases with passage of time as the concentration of reactants decreases.
- The rate of a reaction is same at any time during the reaction
- The rate of a reaction is independent of temperature change.
- The rate of a reaction decreases with increase in concentration of reactant(s)

Brief:



During the course of reaction

- Concentration of reactant is disappearing while that of product appears.
- Rate of reaction become decline as the more reactant has consumed.



217. How does a catalyst increase the rate of a chemical reaction?

- A. By changing the magnitude of the enthalpy change.
- B. By providing a pathway for the reaction with lower E_a .
- C. By increasing the potential energy of the reactants
- D. By giving energy to the reactants so more of them have energies greater than the activation energy

Brief:

- A catalyst increases the rate of reaction by lowering the activation energy.
- A catalyst increases the rate of reaction in both forward and backward directions by providing an alternate pathway with lower activation energy.
- If the activation energy is reduced, more reactants can cross the energy barrier easily.

UNIT FIVE

5. CHEMICAL EQUILIBRIUM

Unit Out Comes

At the end of this section, you will be able to:

- ❖ *Differentiating reversible and irreversible reaction*
- ❖ *State factors affecting chemical equilibrium with regarding to Le-Chatliers principle*
- ❖ *Calculate values for equilibrium constant involving concentration*
- ❖ *State the relationship of K_{eq} to the relative amounts of products and reactants in a given reaction*
- ❖ *Write the equilibrium constant expression for chemical reactions that involve partial pressure*
- ❖ *Show the relationship between K_C and K_P*
- ❖ *Use the equilibrium quotient to predict the direction of the reaction and the position of equilibrium*
- ❖ *Develop skills of calculate equilibrium concentrations from the given initial concentrations*
- ❖ *Determine whether the reactants or products are favored in a chemical reaction*
- ❖ *Use Le-Chatliers principle to explain the effect of changes in temperature, pressure, concentration and presence of catalyst on a reaction*
- ❖ *State the effect of changes in concentration, pressure/volume and temperature on K_{eq}*
- ❖ *Perform an activity to demonstrate the effect of changes in concentration on the position of equilibrium.*
- ❖ *Perform an activity to demonstrate the effect of changes in temperature*
- ❖ *Explain how Le-Chatliers principle is applied in the Haber process (production of NH_3) and in the Contact process (production of H_2SO_4).*

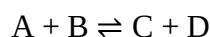
5.1. INTRODUCTION

5.1.1. Reversible and Irreversible reaction

Reversible reactions

The reactions in which the products formed during the reaction combine to give back the reactants are called reversible reactions.

It is indicated by writing two half headed arrow marks between the reactants and product ($\rightleftharpoons$)



Conditions for a reversible reaction:

- The reaction should be carried out in a closed vessel.
- The products should not be removed from the vessel.
- Temperature and pressure should be kept constant.

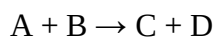
Examples for reversible reactions

- $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
- $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
- $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$

The reactions in which products formed do not combine to give back the reactants are called

Irreversible reactions.

- ✓ The irreversible reactions proceed only in one direction.
- ✓ They are indicated by writing an arrow between reactants and products as:



An irreversible reaction cannot be made reversible even if it is carried out in a closed vessel.

Examples for irreversible reactions:

- $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2 \uparrow$
- $Mg + 2HCl \rightarrow MgCl_2 + H_2 \uparrow$
- $AgNO_3 + KCl \rightarrow AgCl \downarrow + KNO_3$

5.1.2. Attainment and characteristics of chemical equilibrium

Equilibrium is a state in which there are no observable changes as time goes by.

Question: what is dynamic chemical equilibrium?

Answer: any reaction will be in dynamic equilibrium if it's reversible and the rates of the forward and reverse reactions are equal.

An example of dynamic equilibrium is: $\text{NO}_2(\text{g}) + \text{CO}(\text{g}) \rightleftharpoons \text{NO}(\text{g}) + \text{CO}_2(\text{g})$.

Nitrogen dioxide (NO_2) reacts with carbon monoxide (CO) to form nitrogen oxide (NO) and carbon dioxide (CO_2), and, in the reverse reaction, nitrogen oxide and carbon dioxide react to form nitrogen dioxide and carbon monoxide.

Another example of dynamic equilibrium is, say that you prepare a solution that is saturated with an aqueous solution of NaCl . If you then add solid crystals of NaCl , the NaCl will be simultaneously dissolving and recrystallizing within the solution.

The process, $\text{NaCl}(\text{s}) \rightleftharpoons \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$, will be in dynamic equilibrium when the rate of the dissolution of the NaCl equals the rate of recrystallization.

Below is a table showing the key differences between dynamic and static equilibrium.

Dynamic Equilibrium	Static Equilibrium
Reversible	Irreversible
Reaction is still occurring	Reaction has stopped
Rate of forward rxn = Rate of reverse rxn	Both rxn rates are zero
Occurs in a closed system	Can occur in an open or closed system

5.1.3. Chemical equilibrium

When a chemical reaction has reached the equilibrium state, the concentrations of reactants and products remain constant over time, and there are no visible changes in the system. However, there is much activity at the molecular level because reactant molecules continue to form product molecules while product molecules react to yield reactant molecules, it is a dynamic situation.

Chemical equilibrium is the state of the reaction when the macroscopic properties like temperature, pressure, volume and concentration of the reaction do not change with time.

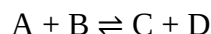
5.1.4. Conditions for attainment of chemical equilibrium

Analogous to the physical systems chemical reactions also attain a state of equilibrium.

- The reactions can occur both in forward and backward directions.
- The rates of the forward and reverse reactions become equal
- The concentrations of the reactants and the products remain constant.
- The equilibrium is dynamic in nature as it consists of a forward reaction in which the reactants give product(s) and reverse reaction in which product(s) gives the original

reactants

For a better comprehension, let us consider a general case of a reversible reaction,



With passage of time, there is accumulation of the products C and D and depletion of the reactants A and B.

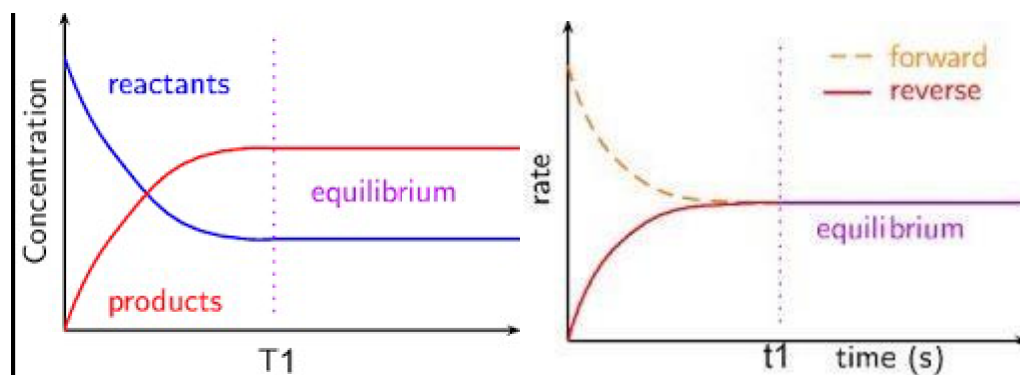


Fig. The concentration vs time graph of the reaction progress

This leads to a decrease in the rate of forward reaction and an increase in the rate of the reverse reaction, eventually, the two reactions occur at the same rate and the system reaches a state of equilibrium. We know from "Factors affecting reaction rates" the rate of the forward reaction is directly proportional to the concentration of the reactants. In other words, as the concentration of the reactants increases, so does the rate of the forward reaction. This can be shown using the following equation:

- ❖ Rate of forward reaction $\propto [A][B]$ or Rate of forward reaction $= k_1[A][B]$.
- ❖ Rate of the reverse reaction is directly proportional to the concentration of the products.
- ❖ Rate of reverse reaction $\propto [C][D]$ or Rate of reverse reaction $= k_2[C][D]$

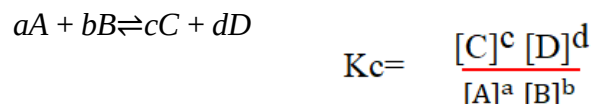
At equilibrium, the rate of the forward reaction is equal to the rate of the reverse reaction. This can be shown using the following equation:

$$k_1 [A][B] = k_2 [C][D] \text{ or } \frac{k_1}{k_2} = \frac{[C][D]}{[A][B]}$$

or, if the constants k_1 and k_2 are simplified to a single constant, the equation becomes:

$$K_c = \frac{[C][D]}{[A][B]}$$

A more general form of the equation for a reaction at chemical equilibrium is:



It is important to note that if a reactant or a product in a chemical reaction is in either the liquid or solid phase, the concentration stays constant during the reaction. Therefore, these values can be left out of the equation to calculate K_c .

For example, in the following reaction: $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$

$$K_c = \frac{[CO][H_2]}{[H_2O]}$$

5.1.5. Characteristics of chemical equilibrium

- Equilibrium is achieved in a closed system. A closed system prevents exchange of matter with the surroundings, so equilibrium is achieved where both reactants and products can react and recombine with each other.
- The rate of the forward reaction is equal to the rate of the backward reaction.
- Equilibrium is dynamic. The reaction hasn't stopped but both forward and backward reactions are still occurring.
- The concentrations of reactants and products remain constant at equilibrium. They are being produced and destroyed at an equal rate.
- At equilibrium there's no change in macroscopic properties.
- Equilibrium can be reached from either direction.
- The free energy change of the reaction is zero at equilibrium

❖ **Derive the relationship between K_c and K_p** (self-practice)

5.1.6. Le Chatelier's principles,

Unless the reaction parameters (concentration, Temperature, pressure/volume) are remained when a stress is applied to a system at equilibrium, the Equilibrium position will shift in the direction that relieves the applied stress.

Stresses include:

- Changes in concentration or pressure of reactants or products by the addition or removal
- Changes in pressure as a result of changes in volume of container
- Changes in temperature (value of K will change)
- Addition of a catalyst

A. Effect of Changes in Concentration on Equilibrium Position

In general, when equilibrium is disturbed by the addition/removal of any reactant/products, Le Chatelier's principle predicts that:

- ❖ The concentration stress of an added reactant/product is relieved by net reaction in the direction that consumes the added substance
- ❖ The concentration stress of a removed reactant/product is relieved by net reaction in the direction that replenishes the removed substance.
- ❖ $Q_c > K_c$, a net reverse reaction (equilibrium shifted to the left)

B. Effect of Changing Pressure on Equilibrium Position

- ✚ A pressure change obtained by changing the volume can affect the yield of products in case of a gaseous reaction where the total number of moles of gaseous reactants and total number of moles of gaseous products are different.
- ✚ In applying Le Chatelier's principle to a heterogeneous equilibrium the effect of pressure changes on solids and liquids can be ignored because the volume (and concentration) of a solution/liquid is nearly independent of pressure.
- ✚ The direction in which the reaction goes to re-establish equilibrium can be predicted by applying the Le Chatelier's principle.
Since pressure has doubled, the equilibrium now shifts in the forward direction, a direction in which the number of moles of the gas or pressure decreases (we know pressure is proportional to moles of the gas).
- ✚ When volume of the reaction mixture is halved, the partial pressure and the concentration are doubled.
- ✚ We obtain the reaction quotient by replacing each equilibrium concentration by double its value.
- ✚ As $Q_c < K_c$, the reaction proceeds in the forward direction

C. Effect of Inert Gas Addition

If the volume is kept constant and an inert gas such as argon is added which does not take part in the reaction, the equilibrium remains undisturbed. It is because the addition of an inert gas at constant volume does not change the partial pressures or the molar concentrations of the substance involved in the reaction. The reaction quotient changes only if the added gas is a reactant or product involved in the reaction.

D. Effect of Temperature Change

Whenever equilibrium is disturbed by a change in the concentration, pressure or volume, the composition of the equilibrium mixture changes because the reaction quotient, Q_c no longer equals the equilibrium constant, K_c . However, when a change in temperature occurs, the value of equilibrium constant, K_c is changed. The temperature dependence of the equilibrium constant depends on the sign of ΔH for the reaction.

- ✓ The equilibrium constant for an exothermic reaction (negative ΔH) decreases as the temperature increases
- ✓ The equilibrium constant for an endothermic reaction (positive ΔH) increases as the temperature increases.

According to Le Chatelier's principle,

- Raising the temperature shifts the equilibrium to left and decreases the equilibrium concentration of ammonia.
- low temperature is favorable for high yield of ammonia, but practically very low temperatures slow down the reaction and thus a catalyst is used

E. Effect of a Catalyst

- A catalyst enhances the rate of a reaction by lowering the reaction's activation energy.
- A catalyst lowers the activation energy of the forward reaction and the reverse reaction to the same extent.
- The presence of a catalyst does not alter the equilibrium constant, nor does it shift the position of an equilibrium system.
- Adding a catalyst to a reaction mixture that is not at equilibrium will simply cause the mixture to reach equilibrium sooner
- The same equilibrium mixture could be obtained without the catalyst, but we might have to wait much longer for it to happen.
- Catalysts lower the activation energy of reactions and make the reaction faster but do not change the equilibrium constant (K_{eq}).

5.1.7. Chemical equilibrium and industry

A. Haber process for the manufacture of ammonia

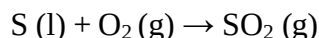


$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}); \Delta H = -92 \text{ KJ/mol}$		
Typical conditions	Pressure	20000 Kpa (200 atm)
	Temperature	380 – 450 °C
	Catalyst	Iron
Equilibrium theory Favors	Low temperature	Exothermic reaction-higher yield at low T
	High pressure	Decrease in number of gaseous molecules
Kinetic theory favors	High temperature	Greater average energy + more frequent collision
	High pressure	More frequent collisions for gaseous molecules
	Catalyst	Lower activation energy
Compromise conditions	Which is better?	A low yield in a shorten time or A high yield over a longer period

B. Contact process for the manufacture of sulphuric acid

✓ Contact process is the industrial manufacturing of sulphuric acid

- Liquid sulphur is sprayed with excess dry air at atmospheric pressure.



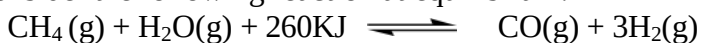
- The oxidation of SO_2 to SO_3 is the most difficult part of making sulfuric acid. The reaction is as follows:



This reaction is an exothermic reaction (energy is released), so by Le Chatelier's Principle, higher temperatures will force the equilibrium position to shift to the left of the equation favoring the production of sulphur dioxide. Lower temperatures would favour the production of the product sulphur trioxide and result in a higher yield. However, the rate of reaching equilibrium at the lower temperatures is extremely low. Higher temperature means equilibrium is established more rapidly but the yield of sulphur trioxide is lower. A temperature of 450°C is a compromise whereby a faster reaction rate results in a slightly lower yield

5.2. Questions with explanation

218. Consider the following reaction at equilibrium:



Which of these is **NOT** a true effect of the stress mentioned?

- A. Rising temperature favors the products formation.
- B. Removal of H_2 decreases the amount of H_2O .
- C. Adding argon gas favors the formation of methane.
- D. Adding CO decreases the amount of H_2 .

Perturbation Effect on Equilibrium Position Effect on K_c

No	Perturbation/Effect	Effect on Equilibrium Position	Effect on K_c
1	Increase in concentration of a reactant, or, decrease in concentration of a single Product.	Shift to the right	None
2	Decrease in concentration of reactant, or, increase in concentration of a single product.	Shift to the left	None
3	Increase the temperature	Shift to endothermic direction	Decrease
4	Decrease the temperature	Shift to exothermic direction	Increase
5	Decreasing the pressure or increasing the volume (for only Gases)	Shift from lower N_0 of mole to higher N_0 of mole	
6	Increasing the pressure or Decreasing the volume (for only Gases)	Shift from higher N_0 mole to lower N_0 of mole	
7	Addition of an inert substance, catalyst, pure liquid, or pure solid	None	None

- ✓ Notice that only a temperature change can affect the value of K_c ; in all other cases the value of K_c remains constant.
- ✓ When the chemical equilibrium shift to the right side:
 - The concentration of product will increase
 - The concentration of reactants will decrease
- ✓ When the chemical equilibrium shift to the left side:
 - The concentration of product will decrease
 - The concentration of reactants will increase

219. In which of these reactions, the equilibrium constant, K_p is Undefined?
- $\text{Ni(s)} + 4\text{CO(g)} \rightleftharpoons \text{Ni(CO)}_4\text{(g)}$
 - $\text{CaCO}_3\text{(s)} \rightleftharpoons \text{CaO(s)} + \text{CO(g)}$
 - $\text{Cr}_2\text{O}_3\text{(s)} + 3\text{CCl}_4\text{(g)} \rightleftharpoons 2\text{CrCl}_3\text{(s)} + 3\text{COCl}_2\text{(g)}$
 - $\text{S(s)} + \text{H}_2\text{SO}_3\text{(aq)} \rightleftharpoons \text{H}_2\text{S}_2\text{O}_3\text{(aq)}$
220. Which one of the following is correct about a dynamic equilibrium? it constant for
- Is a form of static equilibrium
 - Occurs when the rate constant for the forward reaction is equal the rate constant for reverse reaction
 - Only occur in chemical equilibrium
 - Exists when the rate forward reaction is equal to the rate of reverse reaction
221. If the equilibrium concentrations found in the reaction: $2\text{A(g)} + \text{B(g)} \rightleftharpoons \text{C(g)}$ are $[\text{A}] = 2.4 \times 10^{-2} \text{ M}$, $[\text{B}] = 4.6 \times 10^{-3} \text{ M}$ and $[\text{C}] = 6.2 \times 10^{-3} \text{ M}$. Then what is the value of K_c ?
- 1.3×10^{-4}
 - 2.3×10^{-3}
 - 56
 - 5.8×10^2

Brief:

Simple calculation! $K_c = [\text{C}]/[\text{A}]^2[\text{B}]$ substitute the given information

222. The reaction, $\text{CO(g)} + 3\text{H}_2\text{(g)} \rightleftharpoons \text{CH}_4\text{(g)} + \text{H}_2\text{O(g)}$ is in equilibrium at 1300 K in a one litre flask. The gaseous equilibrium mixture contains 0.30 mol of CO, 0.10 mol of H_2 and 0.020 mol of H_2O and an unknown amount of CH_4 in the flask. Determine the concentration of CH_4 in the mixture. The equilibrium constant, K_c for the reaction at the given temperature is 3.90.
- 0.0421
 - 0.0452
 - $2.56 \times 10^{-2} \text{ M}$
 - $5.85 \times 10^{-2} \text{ M}$

Brief:

Let the concentration of methane at equilibrium be x.

- o $\text{CO(g)} + 3\text{H}_2\text{(g)} \rightleftharpoons \text{CH}_4\text{(g)} + \text{H}_2\text{O(g)}$
- o Concentration=Moles/volume
- o concentration at equilibrium: $\text{CO} = 0.3/1 = 0.3$

$$M[\text{H}_2] = 0.1 \text{ mol}/1\text{L} = 0.1 \text{ M}$$

$$[\text{H}_2\text{O}] = 0.02 \text{ mol}/1\text{L} = 0.02 \text{ M}$$

Let concentration of CH_4 be X M

$$\begin{aligned} \frac{[\text{CH}_4\text{(g)}][\text{H}_2\text{O(g)}]}{[\text{CO(g)}][\text{H}_2\text{(g)}]^3} &= K_c \\ \Rightarrow \frac{x \times 0.02}{0.3 \times (0.1)^3} &= 3.90 \\ \Rightarrow x &= \frac{3.90 \times 0.3 \times (0.1)^3}{0.02} \\ &= \frac{0.00117}{0.02} \\ &= 0.0585 \text{ M} \\ &= 5.85 \times 10^{-2} \text{ M} \end{aligned}$$

Hence, the concentration of CH_4 at equilibrium is $5.85 \times 10^{-2} \text{ M}$.

223. If the following reaction is at equilibrium, which one of the following changes will shift the

Equilibrium to the left? $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) + \text{heat}$

- A. Increasing pressure C. adding more N_2 and H_2
 B. Decreasing temperature D. increasing the volume of the reaction container

Brief:

$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) + \text{heat}$		
Increasing pressure	Equilibrium shift to right	Higher to lower No of mole
Decreasing temperature	Equilibrium shift to right	Towards exothermic direction
Adding more N_2 and H_2	Equilibrium shift to right	From higher to lower concentration
Increasing the volume rxn	Equilibrium shift to left	Lower to higher No of mole

224. The equilibrium constant for the reaction.

$2\text{Q} + \text{p} \rightleftharpoons \text{Q}_2\text{P}$ is K_c , then what is the equilibrium constant for the rxn is

$3\text{Q}_2\text{p} \rightleftharpoons 6\text{Q} + 3\text{Y}$

- A. K_c^3 B. $3K_c$ C. $(1/K_c)^3$ D. $(K_c^{-1})^{1/2}$

Brief:

Given the two equations

○ $2\text{Q} + \text{p} \rightleftharpoons \text{Q}_2\text{P} \dots\dots\dots K_c$

○ $3\text{Q}_2\text{p} \rightleftharpoons 6\text{Q} + 3\text{Y} \dots\dots K_c'$

✓ The 2nd equation is obtained by reversing and multiplying by 3 of the 1st equation

✓ The Newly provided K_c , $K_{c_{\text{new}}} = (1/K_c)^{1/3}$

225. For the reaction $\text{MgCO}_3(\text{S}) \rightarrow \text{MgO}(\text{s}) + \text{CO}_2(\text{g})$ K_p is equal to

- A. K_c B. K_cRT C. $K_c(RT)^2$ D. $K_c(RT)^{-1}$

Brief:

✓ K_p and K_c are the equilibrium constant of an ideal gaseous mixture.

✓ K_p is equilibrium constant used when equilibrium concentrations are expressed in atmospheric pressure

✓ K_c is equilibrium constant used when equilibrium concentrations are expressed in molarity.

✓ The process of finding the Reaction Quotient (Q_c) is the same as finding K_c and K_p ,

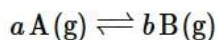
where the products of the reaction is divided by the reactants of the reaction

(*Products/Reactants*) at any time not necessarily at equilibrium.

- $K < Q$: Since there are more products than reactants, the reaction will produce more reactants to reach equilibrium, the reaction favors the reactants.
- $K > Q$: Since there are more reactants than products, the reaction will produce more products to reach equilibrium, the reaction favors the products.
- $K = Q$: There is no change in the products nor reactants, so equilibrium is achieved.

K_c and K_p relationship

For a simple reaction involving gas phase substances



The equilibrium constant can be written in terms of molar concentrations or partial pressures as

$$K_c = \frac{[\text{B}]^b}{[\text{A}]^a} \quad \text{or} \quad K_p = \frac{(P_{\text{B}})^b}{(P_{\text{A}})^a}$$

Assuming that the gas behaves ideally (a common and convenient assumption for general cases), we can write the ideal gas law for each gas such that

$$PV = nRT \quad \rightarrow \quad P_{\text{A}}V = n_{\text{A}}RT \quad \rightarrow \quad P_{\text{A}} = \left(\frac{n_{\text{A}}}{V}\right)RT$$

$$PV = nRT \quad \rightarrow \quad P_{\text{B}}V = n_{\text{B}}RT \quad \rightarrow \quad P_{\text{B}} = \left(\frac{n_{\text{B}}}{V}\right)RT$$

We note that

$$\frac{n_{\text{A}}}{V} \equiv M$$

where M is simply molarity (moles/unit volume). Since $[\text{X}] = M_{\text{x}}$, we get

$$P_{\text{A}} = [\text{A}]RT \quad \text{and} \quad P_{\text{B}} = [\text{B}]RT$$

Substitute these terms into the K_p equation to give

$$K_p = \frac{(P_{\text{B}})^b}{(P_{\text{A}})^a} = \frac{([\text{B}]RT)^b}{([\text{A}]RT)^a}$$

and recognize that

$$\frac{[\text{B}]^b}{[\text{A}]^a} = K_c$$

Plug this relationship into the K_p expression to give

$$K_p = K_c(RT)^{b-a} \quad \rightarrow \quad K_p = K_c(RT)^{\Delta n}$$

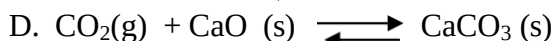
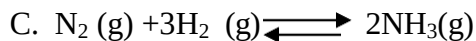
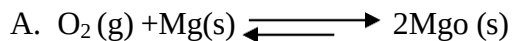
Where Δn is simply the difference of the sum of the stoichiometric coefficients of the products and reactants from a balanced chemical equation like so

$$\Delta n = \sum(\text{products}) - \sum(\text{reactants})$$

It should be clear from this that the only way $K_c = K_p$ within the context of an ideal gas is when $\Delta n = 0$.

$$\begin{aligned} K_p = K_c(RT)^{\Delta n} \quad \rightarrow \quad K_p &= K_c(RT)^0 \\ &= K_c(1) \\ &= K_c \end{aligned}$$

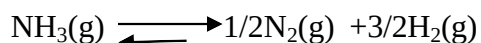
226. Assuming that the following reactions are occurred at room temperature. Identified a reaction which is homogenous



Brief:

- Depends on the physical state of reactant and product, there are two kinds of chemical reactions.
- Homogenous reactions.** Are reaction where all reagents and products are in the same phase is considered homogenous reactions. This means that these reactions can be all solid, all liquid, all gas, or all aqueous but not solids and liquids. E.g: $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$
- Heterogeneous reactions:** are reactions where reagents and products contain at least two different phases. Heterogeneous mixtures have visually distinguishable components, while homogeneous mixtures appear uniform throughout.

227. For the reaction $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ at 400K $K_c = 41$ then what is the value of K_p at the same temperature for the reaction.



A. 0.024

B. 0.16

C. 6.4

D. 5.25

Brief:

- First determine the value of K_c for the reaction: $\text{NH}_3(\text{g}) \rightleftharpoons 1/2\text{N}_2(\text{g}) + 3/2\text{H}_2(\text{g})$
- $K_{c_{\text{new}}} = (1/K_c)^{1/2} = (1/41)^{1/2} = 0.15$
- $K_p = K_c (\text{RT})^{\Delta n}$

228. If the equilibrium is established by initially adding 0.1 mol each of A and B to a 1L container,



then which of the following must be true once the mixture achieves equilibrium?

A. $[\text{A}] = [\text{B}]$ B. $[\text{A}] = [\text{B}]$ C. $[\text{A}] = [\text{B}] = [\text{C}]$ D. $[\text{A}] < [\text{B}]$ E. $[\text{B}] = 2[\text{C}]$

Brief:

In this **equilibrium reaction**: $\text{A} + 2\text{B} \rightleftharpoons 2\text{C}$, $K_c = 320$.

The expression for the equilibrium constant K can be written as:

$$K = [\text{C}]^2 / ([\text{A}] * [\text{B}]^2)$$

Where $[\text{A}]$, $[\text{B}]$, and $[\text{C}]$ represent the molar concentrations of A, B, and C at equilibrium

Given that initially, 0.10 mol each of A and B are added to a 1L container, we can calculate their molar concentrations as follows:

$$[\text{A}] = 0.10 \text{ mol} / 1 \text{ L} = 0.10 \text{ M} \quad [\text{B}] = 0.10 \text{ mol} / 1 \text{ L} = 0.10 \text{ M}$$

Now, let's calculate the **equilibrium** concentration of C using the equilibrium constant K :

$$320 = [C]^2 / (0.10 * 0.10^2)$$

Solving for [C]:

$$320 = [C]^2 / 0.001$$

$$[C]^2 = 320 * 0.001$$

$$[C]^2 = 0.32$$

$$[C] \approx \sqrt{0.32}$$

$$[C] \approx 0.57 \text{ M}$$

So, at equilibrium, the molar concentration of C is approximately 0.57 M.

Now, comparing the molar concentrations of A, B, and C:

$$[A] = 0.10 \text{ M} \quad [B] = 0.10 \text{ M} \quad [C] \approx 0.57 \text{ M}$$

. The [A] and [B] are equal but the concentration of C is higher than that of A and B.

229. Given a K value of 0.43 for the following aqueous equilibrium, suppose sample Z is placed in to water such that its original concentration is 0.033M.

Assume there was zero initial concentration of either A or B. Once equilibrium has occurred, what will be the equilibrium concentration of Z? (self-practice)

$$2A(aq) + B(aq) \rightleftharpoons 2Z(aq) \quad K=0.43$$

A. More than 0.033M

B. Less than 0.033M

B. C.0.033M exactly

D. there is no way to tell

230. When heated at high T^0 , iodine vapor dissociates as follows

$I_2(g) \rightleftharpoons 2I(g)$. In the experiment, a chemist finds that when 0.054 mol of I_2 was placed in a flask of value 0.84L at 578K, the degree of dissociation (is the fraction of I_2 dissociated) was of 0.025. What is K_P and K_C respectively? (self-practice)

A. 7.8×10^{-3} and 1.6×10^{-4}

B. 1.83 and 0.038

E. None

C. 1.2×10^{-13} and 2.53×10^{-15}

D. 4.24 and 0.0895

231. Which of the following is false?

A. The smaller the activation energy, the slower a reaction tends to be.

B. In general, reaction rates increase with increase in temperature.

C. The heat of reaction is the difference between the energy of the products and the energy of the reactants.

D. All of the above

Brief:

Facts about rate of reaction

- ✓ The activation energy of a chemical reaction is closely related to its rate. Specifically, the higher the activation energy, the slower the chemical reaction will be. This is because molecules can only complete the reaction once they have reached the top of the activation energy barrier.
- ✓ Increasing the temperature increases, the average kinetic energy of molecules also increase. As more molecules move faster, the number of molecules moving fast enough to react increases, which results in faster formation of products
- ✓ The heat of reaction is the difference between the energy of the products and the energy of the reactants::

232. At equilibrium_____.

- A. The rates of the forward and reverse reactions are equal
- B. The rate constants of the forward and reverse reactions are equal
- C. All chemical reactions have stopped
- D. The value of the equilibrium constant is very large

Brief:

Characteristics of Chemical Equilibrium:

- Chemical equilibrium is dynamic (the reaction never stopped) in nature
- Forward and reverse reactions have the same rate in chemical equilibrium.
- At equilibrium, all of the reactants and products are present.
- A catalyst has no effect on the equilibrium position.
- At equilibrium, the concentrations of the reactants and products are constant.

233. $2\text{NO}(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2\text{NOCl}(\text{g})$, $\Delta H = -78\text{KJ}$ what condition of temperature and pressure will produce the highest yield of NOCl at equilibrium ?

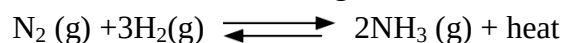
- | | |
|-------------------------------------|------------------------------------|
| A. High temperature , high pressure | C. high temperature , low pressure |
| B. Low temperature , high pressure | D. low temperature , low pressure |

Brief:

- Optimum conditions are the condition that favors the equilibrium to the desired product.
- The optimum condition for the reaction: $2\text{NO}(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2\text{NOCl}(\text{g})$, $\Delta H = -78\text{KJ}$ are:
 - Increasing concentration of reactants ($2\text{NO}(\text{g})$ or $\text{Cl}_2(\text{g})$)
 - Decreasing the Temperature

- Decreasing volume
- Increasing Pressure

234. Which of the following statement is not true about the reaction given below?



- A. The equilibrium constant, $K_C = K_P (RT)^2$
- B. High pressure and low temperature increase yields of ammonia
- C. Addition of temperature favour & ammonia production
- D. Decrease in concentration of nitrogen shifts the equilibrium to left

235. When $K_c \gg 1$ for a chemical reaction:

- A. The equilibrium would be achieved rapidly.
- B. The equilibrium would be achieved slowly.
- C. Reactants would be much more stable than products.
- D. Product concentrations would be much greater than reactant concentrations at equilibrium
- E. Reactant concentrations would be much greater than product concentrations at equilibrium.

Brief:

⇒ If the K_c value is large ($K_c \gg 1$), the equilibrium lies to the right and the reaction mixture contains mostly products.

⇒ If the K_c value is small ($K_c \ll 1$), the equilibrium lies to the left and the reaction mixture contains mostly reactants.

236. What would be the effect of decreasing the pressure by increasing the volume on the following system at equilibrium? $2\text{CO}(\text{g}) + \text{O}_2 \rightleftharpoons 2\text{CO}_2(\text{g})$

- A. The K_p value would get smaller.
- B. The K_p value would get constant.
- C. The equilibrium would be perturbed and would show a net shift to the left.
- D. The equilibrium would be perturbed and would show a net shift to the right.

Brief:

For the given reaction: $2\text{CO}(\text{g}) + \text{O}_2 \rightleftharpoons 2\text{CO}_2(\text{g})$

- On decreasing the pressure/increasing the volume of the reaction vessel
 - Equilibrium shift from right ($n_p=2$) to left ($n_r=3$)
 - The concentration of reactant increase while product decrease
 - The value of K_p becomes smaller.

237. The reaction: $2\text{NO}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{O}_2(\text{g})$ is endothermic. If additional O_2 is added to a vessel at equilibrium what will ultimately happen to the NO_2 concentration? It will (self-practice)

- A. increase
- B. Decrease
- C. remain the same
- D. none

238. If the equilibrium constant for the reaction $\text{N}_2 + 3\text{Cl}_2 \rightleftharpoons 2\text{NCl}_3$ is K_p

$= 1.0 \times 10^3$, what is expected to happen to this mixture? (self-practice)

- A. It is not in equilibrium and will proceed in the forward direction.
- B. It is not in equilibrium and will proceed in the reverse direction.
- C. The mixture is in equilibrium and will not change.
- D. The reaction attain equilibrium

239. In which of the following systems would the number of moles of the substances present at equilibrium NOT be shifted by a change in the volume of the system at constant temperature?

- A. $\text{CO(g)} + \text{NO(g)} \rightleftharpoons \text{CO}_2\text{(g)} + \frac{1}{2} \text{N}_2\text{(g)}$
- B. $\text{N}_2\text{(g)} + 3 \text{H}_2\text{(g)} \rightleftharpoons 2 \text{NH}_3\text{(g)}$
- C. $\text{N}_2\text{(g)} + 2 \text{O}_2\text{(g)} \rightleftharpoons 2 \text{NO}_2\text{(g)}$
- D. $\text{NO(g)} + \text{O}_3\text{(g)} \rightleftharpoons \text{NO}_2\text{(g)} + \text{O}_2\text{(g)}$

240. In the reaction $\text{X} + \text{Y} \rightleftharpoons \text{Z} + \text{W}$, if we start equal concentration of X & Y, at equilibrium we find that the mole of Z is twice of X. what is the equilibrium constant of the reaction?

- A. 4
- B. 2
- C. $\frac{1}{4}$
- D. $\frac{1}{2}$

5.3. Self-practice Questions

241. Which of the following best describes a system at equilibrium?
- A. Forward reaction stops completely
 - B. Concentrations of reactants and products are equal
 - C. Rate of forward reaction equals the rate of reverse reaction
 - D. Equilibrium constant is always 1
242. Which factor affects the equilibrium constant K?
- A. Pressure
 - B. Concentration
 - C. Temperature
 - D. Catalyst
243. For a reaction with $K_c > 1$, what can be concluded?
- A. Reactants are favored
 - B. Products are favored
 - C. System is at equilibrium
 - D. Rate of reaction is very fast
244. If the reaction quotient Q is greater than K, the reaction will shift:
- A. Toward reactants
 - B. Toward products
 - C. No change
 - D. Cannot determine
245. Adding a catalyst to a reaction at equilibrium will:
- A. Shift equilibrium right
 - B. Shift equilibrium left
 - C. Equilibrium is reached faster
 - D. Change the equilibrium constant
246. Which statement is true about dynamic equilibrium?
- A. The reaction stops
 - B. Forward and reverse reactions occur at equal rates

- C. Reactant concentration becomes zero
D. More products than reactants
247. Adding an inert gas at constant volume affects equilibrium how?
A. No effect
B. Shift toward reactants
C. Shift toward products
D. Changes K
248. What is the unit of K_c for $A + B \rightleftharpoons CA + B$
A. M^{-1} B. M C. No units D. Depends on temperature
249. Increasing temperature in an exothermic reaction shifts equilibrium:
A. Left B. Right C. No effect D. K remains the same
250. Decreasing temperature in an endothermic reaction:
A. Shifts right B. Shifts left C. Increases K D. No effect
251. For the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$
A. Shift right B. Shift left C. No effect D. Change K
252. Which change affects both the position and value of K?
A. Concentration B. Pressure C. Catalyst D. Temperature
253. What happens when a reaction at equilibrium is diluted?
A. No effect B. Shifts toward more particles
C. Shifts toward fewer particles D. K changes
254. For a reaction at equilibrium, what does a very small K indicate?
A. More products B. More reactants C. Equal amounts D. Fast reaction
255. How does K_p relate to K_c ?
A. $K_p = K_c(RT)^{\Delta n}$ B. $K_p = K_c/RT$ C. $K_p = K_c$ D. $K_p = (RT)K_c$

UNIT SIX

6. ORGANIC COMPOUNDS

Unit Outcomes

At the end of this unit, students will be able to

- ❖ *Explain the difference between organic and inorganic compounds*
- ❖ *State the classification of organic compound*
- ❖ *Writing the IUPAC Name of alkanes, alkenes and alkynes having different structure*
- ❖ *Comparing the properties of hydrocarbon, Alkane, Alkene, Alkyne and derivative of organic compound like alcohol carboxylic acid, ester*
- ❖ *Explain the unsaturated test of hydrocarbon: Bayer reagent and bromine water.*
- ❖ *Compare and contrast the physical properties of alcohol types (primary, secondary, tertiary)
Or (monohydric, dihydric, trihydric)*

6.1. INTRODUCTION

6.1.1. Hydrocarbon

➤ Alkanes

- ✓ Alkanes are saturated hydrocarbons with the general formula C_nH_{2n+2} .
- ✓ Named using the IUPAC system, where the longest carbon chain is identified, and substituents are numbered with the lowest locants.
- ✓ Examples: $CH_4 \rightarrow$ Methane, $C_2H_6 \rightarrow$ Ethane $C_3H_8 \rightarrow$ Propane
- ✓ Alkanes have different physical properties as given below in the table.

Property	Characteristic
State	Gases (C_1-C_4), Liquids (C_5-C_{17}), Solids ($>C_{18}$)
Solubility	Insoluble in water, soluble in non-polar solvents
Boiling Point	Increases with molecular weight
Reactivity	Low, undergoes combustion
Combustion	Produces CO_2 and H_2O

- ✓ They are prepared by the following methods.

- * Hydrogenation of Alkenes: $Alkene + H_2 \rightarrow Alkane$
- * Wurtz Reaction: $R-X + 2Na + R'-X \rightarrow R-R' + 2NaX$
- * Decarboxylation: $RCOONa + NaOH \rightarrow RH + Na_2CO_3$

➤ Alkenes

- ✓ Alkenes are unsaturated hydrocarbons containing at least one $C=C$ double bond, with the general formula C_nH_{2n} .
- ✓ Named by replacing "-ane" with "-ene," with numbering prioritizing the double bond position
- ✓ Examples: $C_2H_4 \rightarrow$ Ethene, $C_3H_6 \rightarrow$ Propene
- ✓ Alkenes have different physical properties as given below in the table.

Property	Characteristic
Polarity	Non-polar/slightly polar
Boiling Point	Lower than alkanes due to weaker van der Waals forces
Reactivity	Higher than alkanes (π -bond reactivity)
Isomerism	Shows cis-trans isomerism

- ✓ They are prepared by the following methods.

- * Dehydration of Alcohols: $R-OH \rightarrow Alkene + H_2O$
- * Dehydrohalogenation: $R-X + Base \rightarrow Alkene + HX$
- * Cracking of Alkanes: $C_nH_{2n+2} \rightarrow Alkenes + Alkanes$

➤ Alkynes

- ✓ Alkynes are hydrocarbons containing at least one $\text{C}\equiv\text{C}$ triple bond, with the general formula $\text{C}_n\text{H}_{2n-2}$.
- ✓ Named by replacing "-ane" with "-yne."
- ✓ Examples: $\text{C}_2\text{H}_2 \rightarrow$ Ethyne (Acetylene), $\text{C}_3\text{H}_4 \rightarrow$ Propyne
- ✓ Alkynes have different physical properties as given below in the table.

Property	Characteristic
Bonding	One sigma and two pi bonds
State	Lower alkynes are gases, higher ones are liquids/solids
Reactivity	More reactive than alkanes/alkenes
Solubility	Insoluble in water, soluble in organic solvents

- ✓ They are prepared by the following methods
 - * Dehydrohalogenation of Dihalides: $\text{VicinalDihalide} + \text{Base} \rightarrow \text{Alkyne}$
 - * Kolbe Electrolysis: $\text{Sodium Carboxylate} \rightarrow \text{Alkyne} + \text{CO}_2$

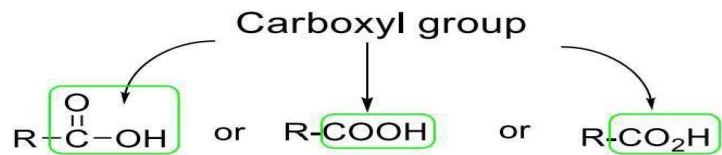
6.1.2. Benzene

- ✓ Aromatic hydrocarbon, formula C_6H_6 .
- ✓ Planar, cyclic, conjugated π -system follows Hückel's rule.
- ✓ Highly stable, does not undergo addition reactions easily.
- ✓ Undergoes the following substitution reactions:
 - Nitration: $\text{Benzene} + \text{HNO}_3 \rightarrow \text{Nitrobenzene}$
 - Halogenation: $\text{Benzene} + \text{Cl}_2 \rightarrow \text{Chlorobenzene}$
 - Sulfonation: $\text{Benzene} + \text{H}_2\text{SO}_4 \rightarrow \text{Benzenesulfonic Acid}$
- ✓ Prepared by
 - * Decarboxylation of Aromatic Acids: $\text{C}_6\text{H}_5\text{COONa} + \text{NaOH} \rightarrow \text{C}_6\text{H}_6 + \text{Na}_2\text{CO}_3$
 - * Reduction of Phenol: $\text{C}_6\text{H}_5\text{OH} + \text{Zn} \rightarrow \text{C}_6\text{H}_6 + \text{ZnO}$
 - * Cyclization of Acetylene: $3\text{C}_2\text{H}_2 \rightarrow \text{C}_6\text{H}_6$

6.1.3. Derivative Hydrocarbon

➤ Carboxylic acids

Are organic compounds containing the carboxyl group (-COOH), wherein the hydroxyl group (-OH) is directly attached to the carbonyl (C=O) group.



Three representations of a carboxylic acid

Nomenclature of Carboxylic Acids

The common names of some basic carboxylic acids are derived from Latin names that indicate the first original natural source of the carboxylic acid.

Structure of Acid	Natural Source	Common Name
$\text{H}-\overset{\text{O}}{\underset{\text{ }}{\text{C}}}-\text{OH}$	Ants (Formica)	Formic acid
$\text{CH}_3-\overset{\text{O}}{\underset{\text{ }}{\text{C}}}-\text{OH}$	Vinegar (Acetum)	Acetic acid
$\text{CH}_3\text{CH}_2-\overset{\text{O}}{\underset{\text{ }}{\text{C}}}-\text{OH}$	Basic Fat (Propio)	Propionic acid
$\text{CH}_3\text{CH}_2\text{CH}_2-\overset{\text{O}}{\underset{\text{ }}{\text{C}}}-\text{OH}$	Rancid butter (Butyrum)	Butyric acid
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2-\overset{\text{O}}{\underset{\text{ }}{\text{C}}}-\text{OH}$	Present in a Valerian herb	Valeric acid
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2-\overset{\text{O}}{\underset{\text{ }}{\text{C}}}-\text{OH}$	Goat (Caper)	Caproic acid

Common Names of Carboxylic Acids

The common name of a carboxylic acid (R-COOH) is derived by adding the suffix -ic acid to a prefix representing the chain length of the carboxylic acid.

# of Carbons	Prefix	Common Name of Acid
1	Form-	Formic acid
2	Acet-	Acetic acid
3	Propion-	Propionic acid
4	Butyr-	Butyric acid
5	Valer-	Valeric acid
6	Capro-	Caproic acid
Aromatic acid	Benzo-	Benzoic acid

IUPAC Nomenclature of Aliphatic Carboxylic Acids

IUPAC names of straight chain aliphatic carboxylic acids are derived by adding the suffix –oic acid to the systematic name of the parent hydrocarbon. They are named as alkanoic acids

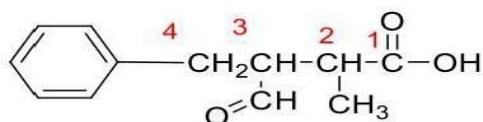
# of Carbons	Structure & IUPAC Name of Alkane	Structure & IUPAC Name of Acid
1	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$ Methane	$\begin{array}{c} \text{O} \\ \\ \text{H}-\text{C}-\text{OH} \end{array}$ Methanoic acid
2	CH_3-CH_3 Ethane	$\begin{array}{c} \text{O} \\ \\ \text{CH}_3-\text{C}-\text{OH} \end{array}$ Ethanoic acid
3	$\text{CH}_3\text{CH}_2-\text{CH}_3$ Propane	$\begin{array}{c} \text{O} \\ \\ \text{CH}_3\text{CH}_2-\text{C}-\text{OH} \end{array}$ Propanoic acid

Systematic Nomenclature of Substituted Carboxylic Acids

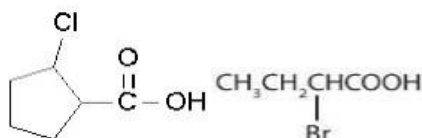
The systematic names of substituted aliphatic carboxylic acids are derived by:

- First identifying the parent chain that contains most, if not all, the carboxyl groups.
- Number the parent chain from the carbon of the carboxyl group i.e the carboxyl carbon is C-1.
- Identify the substituents and assign each substituent a locator/address number (2,3,4...etc.) consistent with the numbering in the parent chain.
- Arrange the names of the substituents in alphabetical order in the systematic name of the poly-substituted carboxylic acid.

Examples:



3-Formyl-2-methyl-4-phenylbutanoic acid



2-chlorocyclopentanoic acid



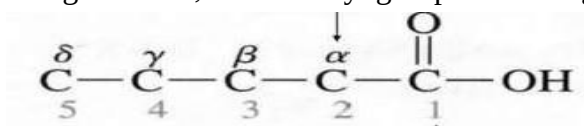
4-chlorobutanoic acid

Naming using Greek letters

There are two ways to identify substituent carbons in carboxylic acid:

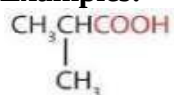
- ✓ numbers or
- ✓ Greek letters.

Using numbers, the carboxyl group carbon is given the number one.

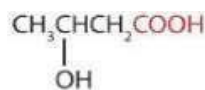


Greek letters are used to designate the position of substituent relative to the carbon of the carboxyl group. The carbon of the carboxyl group is NOT given a Greek letter.

Examples:



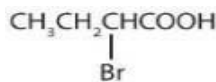
-Methylpropionic acid



-Hydroxybutyric acid



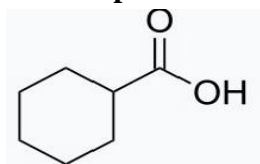
γ -chlorobutyric acid



α -bromobutyric acid

Systematic Nomenclature of Cyclic Carboxylic Acids

The systematic name of a carboxylic acid in which the COOH group is attached directly to a ring is derived by adding a suffix – carboxylic acid to the name of the attached cycloalkane or cycloalkene or arene. **Example:**

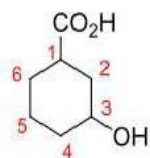


Cyclohexanecarboxylic acid

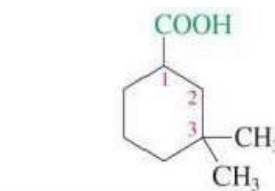
When such carboxylic acids are substituted, the carbon of the COOH group is itself not numbered, but it is, by convention, taken to be attached to C-1 of the ring.

Systematic Nomenclature of Substituted Cyclic Carboxylic Acids

For substituted carboxylic acids with the carboxyl group attached directly to a ring, the carbon of the COOH group is itself not numbered, but it is, by convention, taken to be attached to C-1 of the ring.

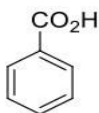


3-Hydroxycyclohexanecarboxylic acid



3,3-dimethylcyclohexanecarboxylic acid

Under this system of nomenclature, benzoic acid would be named as benzenecarboxylic acid.

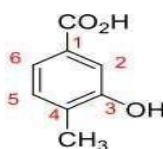


Benzenecarboxylic acid (benzoic acid)

IUPAC Nomenclature of Substituted Aromatic Carboxylic Acids

Substituted aromatic acids with one carboxyl group are named as derivatives of benzoic acid, with the position of substituents being cited using the locators (2,3etc) according to their position on the benzene ring relative to the carboxyl group. The carbon on which the carboxyl group is attached is by convention C-1

Example:



3-Hydroxy-4-methylbenzoic acid or 3-Hydroxy-4-methylbenzenecarboxylic acid

Typically aromatic acids of the form Ar-CO₂H are named as derivatives of benzoic acids, with ortho, meta and para indicating the location relative to the carboxyl group. (Recall that this is non-IUPAC).



Systematic Nomenclature of Dicarboxylic Acids

The systematic name of an open chain aliphatic dicarboxylic acid is derived by adding a suffix -dioic acid to the name of the parent hydrocarbon i.e. alkanedioic acid.

Systematic Name	Common Name	Structure
Ethanedioic acid	Oxalic acid	HO ₂ C-CO ₂ H
Propanedioic acid	Malonic acid	HO ₂ CCH ₂ CO ₂ H
Butanedioic acid	Succinic acid	HO ₂ C(CH ₂) ₂ CO ₂ H
Pentanedioic acid	Glutaric acid	HO ₂ C(CH ₂) ₃ CO ₂ H
Hexanedioic acid	Adipic acid	HO ₂ C(CH ₂) ₄ CO ₂ H
Heptanedioic acid	Pimelic acid	HO ₂ C(CH ₂) ₅ CO ₂ H

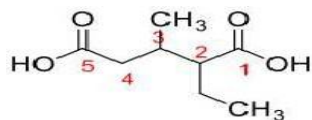
Systematic Nomenclature of Substituted Dicarboxylic Acids

The systematic name of an aliphatic dicarboxylic acid is derived by:

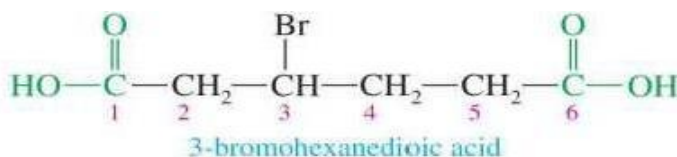
- i. First identifying the parent chain that contains the two carboxylic acid groups and then adding a suffix -dioic acid to the name of the parent hydrocarbon.

- ii. The parent chain is numbered from the end that gives the substituents in the chain the lowest possible address number.
- iii. The substituents are arranged in alphabetical order in the full name of the dicarboxylic acid.

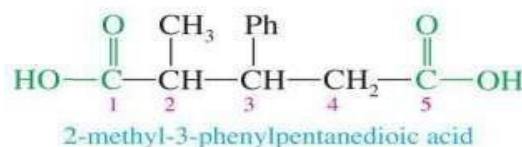
Examples:



2-Ethyl-3-Methylpentanedioic acid



3-bromohexanedioic acid



2-methyl-3-phenylpentanedioic acid

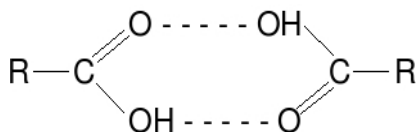
Systematic Nomenclature of Cyclic Dicarboxylic Acids

The systematic name of a cyclic aliphatic or aromatic dicarboxylic acid is derived by adding a suffix -dicarboxylic acid to the name of the parent cycloalkane or arene i.e

cycloalkanedicarboxylic acid or arenedicarboxylic acid.

➤ **Physical Properties of Carboxylic Acids**

- ✓ Carboxylic acids hydrogen bond to themselves to form a dimer:



- ✓ Carboxylic acids also form hydrogen bonds to water molecules:
- ✓ Since carboxylic acids can form more than one set of hydrogen bonds, their boiling points are usually higher than those of other molecules of the same molecular weight (MW).
- ✓ Low-MW carboxylic acids are generally liquids at room temp. (often, they are somewhat oily); higherMW carboxylic acids are generally waxy solids.
- ✓ Carboxylic acids with 12 to 20 carbon atoms are often referred to as fatty acids, since they are found in the triglycerides in fats and oils (more later).
- ✓ Short-chain carboxylic acids are also generally more soluble in water than compounds of similar MW, since they can hydrogen bond to more than one water molecule.
- ✓ As the number of carbons in a carboxylic acid series becomes greater, the boiling point increases and the solubility in water decrease.

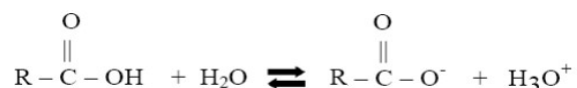
❖ Reactions of Carboxylic Acids

The reactions of carboxylic acids can be directed to various sites on the carboxyl group

Reactions of carboxylic acids can be placed into four categories:

- Reactions at the acidic hydrogen on the carboxyl group.
- Reactions at the carbonyl group
- Reactions at the carboxylate oxygen
- Reactions that lead to loss of the carboxyl group as CO_2

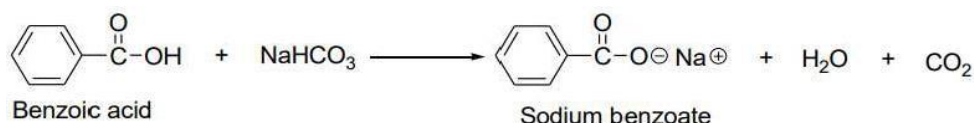
❖ Reaction with water



❖ Reaction of Carboxylic Acids with Sodium Bicarbonate

Most carboxylic acids ($\text{pK}_a \approx 5$) are stronger acids than carbonic acid (H_2CO_3) ($\text{pK}_a 6.4$). Consequently they displace carbonic acid from its salts (hydrogen carbonates). The most reliable test for carboxylic acids employs NaHCO_3 leading to evolution of CO_2 . This is commonly called the bicarbonate test for carboxylic acids.

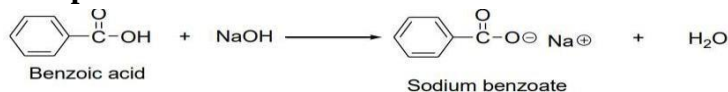
Example:



❖ Reactions of Carboxylic Acids with Strong Bases

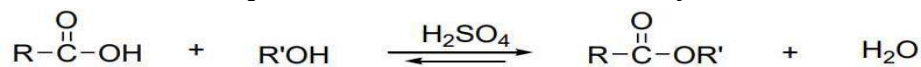
Bases such as metal hydroxides (NaOH and KOH) and amines abstract the acidic proton on carboxylic acids to form carboxylate salts.

Example

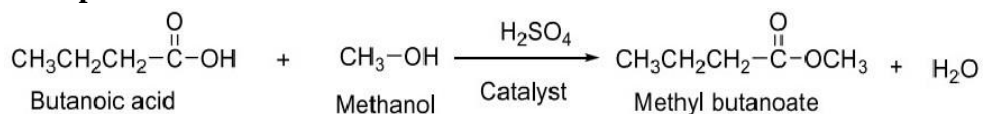


❖ Acid-Catalysed Esterification of Carboxylic Acids

The traditional method for converting carboxylic acids to esters is through an acid-catalyzed esterification in the presence of an alcohol: Commonly referred to as the Fischer esterification.



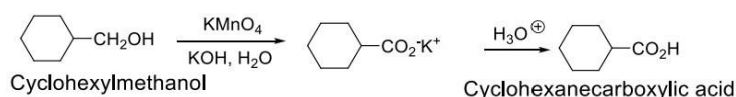
Example:

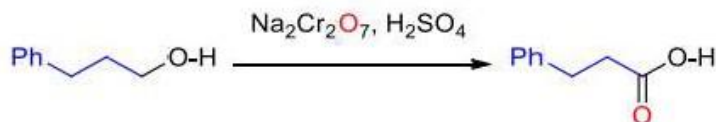


➤ Synthesis of Carboxylic Acids

A. Oxidation of Primary Alcohols to Carboxylic Acids

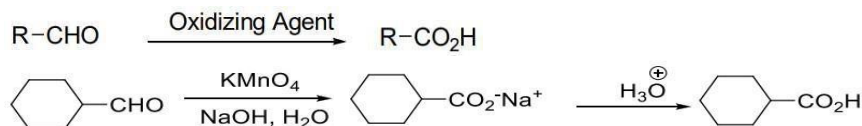
The best conditions for the oxidation of primary alcohols to carboxylic acids is under the basic conditions employing potassium permanganate.





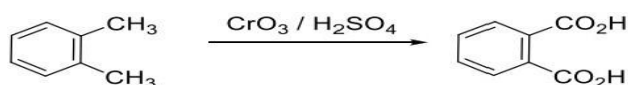
B. Oxidation of Aldehydes to Carboxylic Acids

Aldehydes can be oxidised to carboxylic acids by a variety of oxidizing agents. Both strong and mild oxidizing agents may be employed successfully.



C. Side-Chain Oxidation of Alkylbenzenes to Benzoic Acids

Oxidation of alkylbenzenes using strong oxidizing agents provides benzoic acids. The entire alkyl chain, regardless of its length, is oxidised to a carboxyl (-COOH) group



Fatty acids

- ✓ These are monocarboxylic organic acids, which usually contain an even number of carbon atoms.
- ✓ The general formula of fatty acid is R-COOH
- Example:** $\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{COOH}$
- ✓ They are further classified into: Saturated & Unsaturated fatty acids according to absence or presence of double bonds.

A. Saturated fatty acids

- They have no double bonds.
- They have the general formula: $\text{CH}_3 - (\text{CH}_2)_n - \text{COOH}$.
- They are further classified, according to the number of carbon atoms, into: Short chain & Long chain fatty acids.

The following table shows the formula of the most common saturated fatty acids.

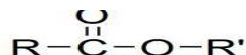
Common name	Formula
Acetic acid (2C)	$\text{CH}_3 - \text{COOH}$
Butyric acid (4C)	$\text{CH}_3 - (\text{CH}_2)_2 - \text{COOH}$
Caproic acid (6C)	$\text{CH}_3 - (\text{CH}_2)_4 - \text{COOH}$
Palmitic acid (16C)	$\text{CH}_3 - (\text{CH}_2)_{14} - \text{COOH}$
Stearic acid (18C)	$\text{CH}_3 - (\text{CH}_2)_{16} - \text{COOH}$
Arachidic acid (20C)	$\text{CH}_3 - (\text{CH}_2)_{18} - \text{COOH}$
Lignoceric acid (24C)	$\text{CH}_3 - (\text{CH}_2)_{22} - \text{COOH}$

B. Unsaturated fatty acids

- They have one or more double bonds.
- Most of the double bonds are of cis type

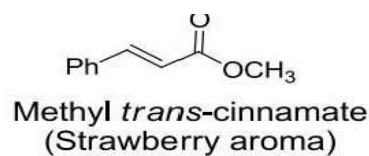
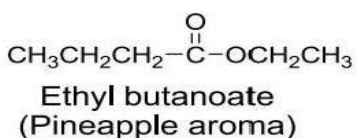
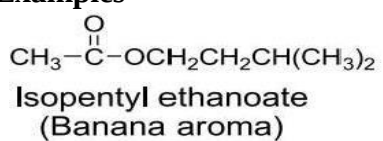
Esters

Structurally, an ester is a compound that has an alkoxy (OR) group attached to the carbonyl group.



R may be H, alkyl or aryl, while R' may be alkyl or aryl only. Esters are widespread in nature. Many of the fragrances of flowers and fruits are due to the esters present. Ethyl butanoate is the chief component that accounts for the pineapple-like aroma and flavour of pineapples.

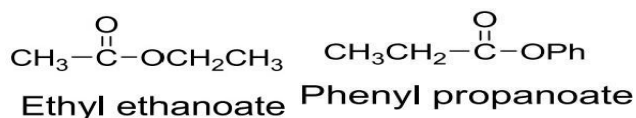
Examples

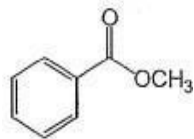
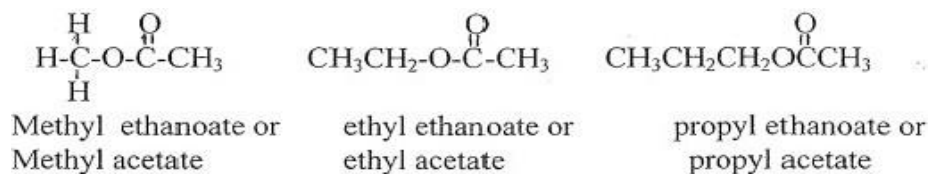


Nomenclature of Esters

- ✚ Names of esters consist of two words that reflect the composite structure of the ester. The first word is derived from the alkyl group of the alcohol component, and the second word from the carboxylate group of the carboxylic acid component of the ester.
- ✚ The name of the carboxylate portion is derived by substituting the -ic acid suffix of the parent carboxylic acid with the -ate suffix.
- ✚ The alkyl group is cited first followed by the carboxylate group separated by a space. An ester is thus named as an alkyl alkanoate.

Examples

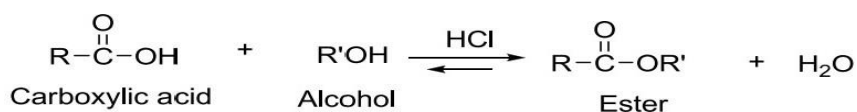




Methyl benzoate

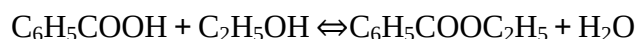
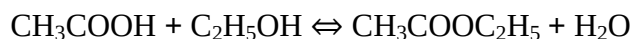
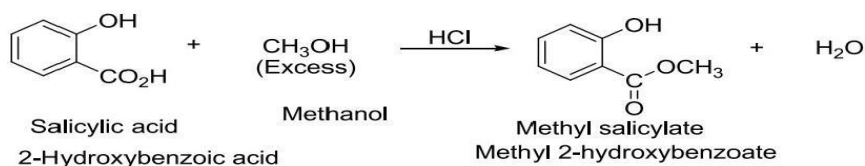
Synthesis of Esters

❖ Acid-Catalysed Esterification of a Carboxylic Acid and an Alcohol



The acid-catalysed reaction of carboxylic acids and alcohols provides esters. Typically, a catalytic amount of a strong inorganic (mineral) acid such as H_2SO_4 , HCl and H_3PO_4 is used.

Examples:



➤ Physical properties

- ❖ Most of the esters are liquids with much lower boiling point than those of the acids or alcohols of nearly equal weight, this is due to the absence of the polar hydroxyl group which found in alcohols and acids and leads to the association of the alcohol and carboxylic acid molecules with hydrogen bonds.
- ❖ Esters have a neutral effect on litmus.
- ❖ The boiling point of the ester is less than the boiling point of the acid and alcohol forming it due to the absence of polar hydroxyl group (presents in alcohols and acids) which has the ability to form hydrogen bonds between molecules and water.

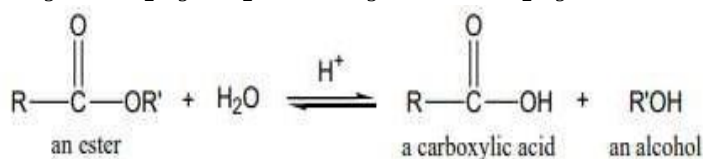
- ❖ The solubility degree of ester in water is less than that of the corresponding acid due to the absence of polar hydroxyl group (presents in alcohols and acids) which has the ability to form hydrogen bonds between molecules and water.

➤ Reactions of Ester

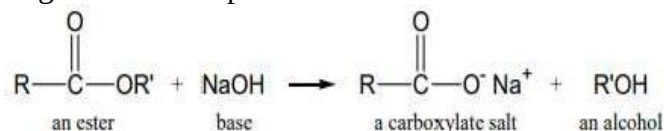
Hydrolysis of esters: Alcohol and acid are produced from the hydrolysis of ester, this reaction reverse to ester formation, Hydrolysis may take place by the use of dilute mineral acids as a catalyst and is called acid hydrolysis, Dilute mineral acid (H^+) is used to prevent the reversible reaction.

This is essentially the reverse reaction of the synthesis of esters from carboxylic acids and alcohols.

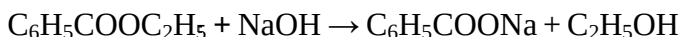
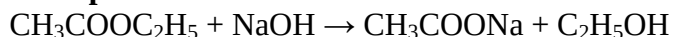
Example: $CH_3COOC_2H_5 + H_2O \rightarrow CH_3COOH + C_2H_5OH$



Hydrolysis of esters may also be carried out by heating with aqueous alkalis, to produce the alcohol and the salt of the acid, this is called alkaline-hydrolysis or saponification, (since soap is the sodium salt of high carboxylic acids), NaOH is added to react with the produced acid converting it to salt and prevent the reversible reaction.

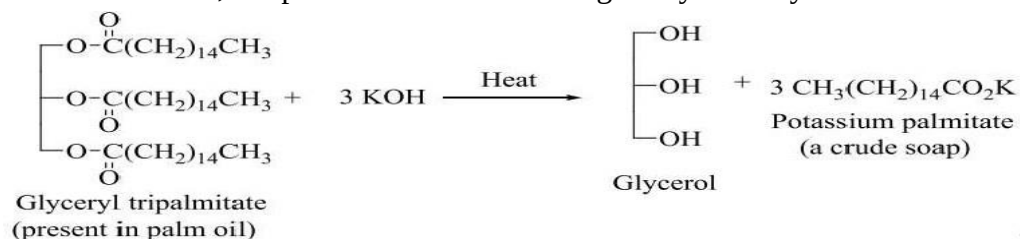


Examples:



For historical reasons, ester hydrolysis in aqueous hydroxide (KOH or NaOH) is called saponification because it was used in the manufacture of soap through the reaction of oils or fats (triesters or triglycerides) with lye (contains mainly KOH).

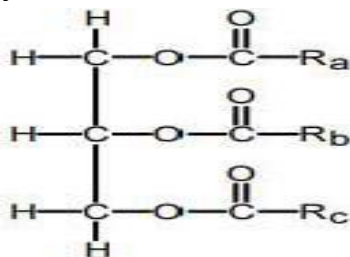
- Hydrolysis of ester is the reaction of the ester with water (in an acidic medium) to form acid and alcohol
- Saponification is the heating of ester with aqueous alkalis to produce the alcohol and the salt of the acid, Soap is the sodium salt of high fatty carboxylic acids.



Fats and Oils

Fats and oils belong to a group of biological substances called lipids. Lipids are biological chemicals that do not dissolve in water. Fats differ from oils only in that they are solid at room

temperature, while oils are liquid. Fats and oils share a common molecular structure, which is represented by the formula below.



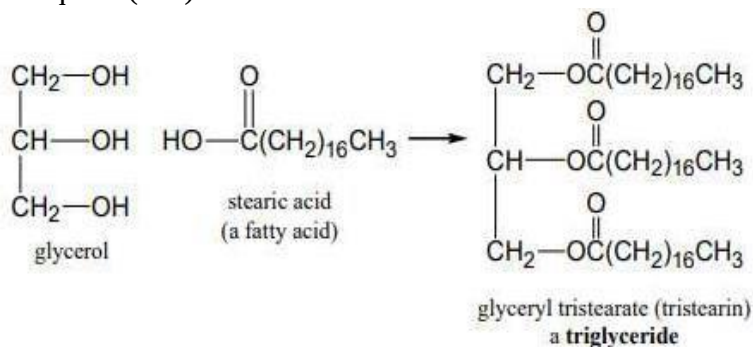
This structural formula shows that fats and oils contain three ester functional groups. Fats and oils are esters of the tri-alcohol, glycerol (or glycerine). Therefore, fats and oils are commonly called triglycerides, although a more accurate name is triacylglycerols. One of the reactions of triglycerides is hydrolysis of the ester groups.

Fats and oils:

- ❖ They are esters of fatty acids with glycerol.
- ❖ They are similar chemically, but they differ in physical properties as oils are liquid while fats are solid at room temperature.
- ❖ They are called triglycerides because they are triesters formed of glycerol and 3 fatty acids.

Triglycerides

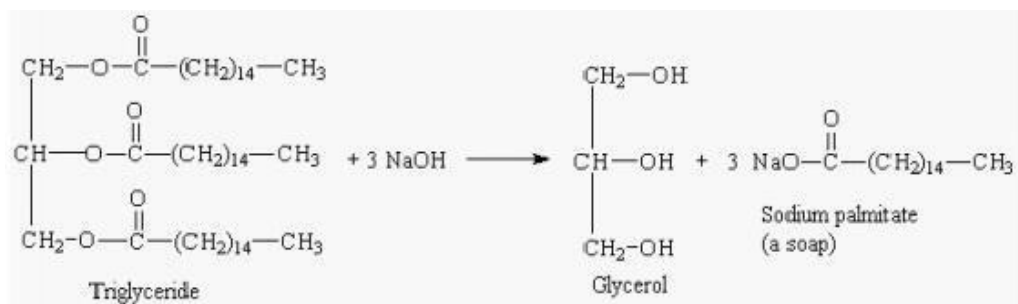
- Fats and oils are both triglycerides; the only difference is that fats are solids at room temperature and oils are liquids at room temperature.
- Long-chain carboxylic acids, often referred to as fatty acids, are stored by living organisms by combining them with glycerol to produce tri-esters called triglycerides.
- Triglycerides at room temperature are usually either solids or semi-solids (fats), or viscous liquids (oils).



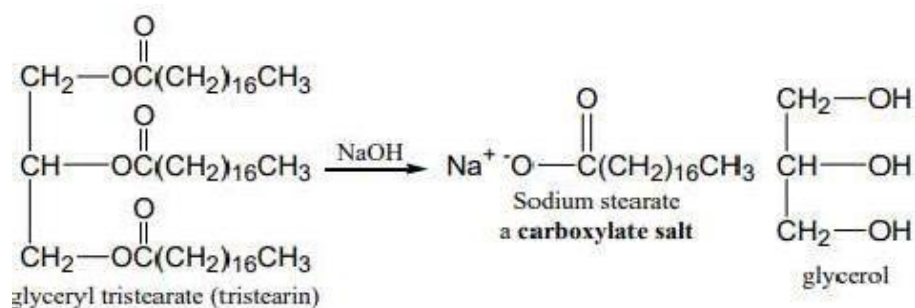
Hydrolysis of triglycerides

The hydrolysis of triglycerides can be carried out with base and this is the common way of doing it on an industrial scale. The hydrolysis and neutralization are carried out simultaneously and produces soap in a one step reaction called **saponification**.

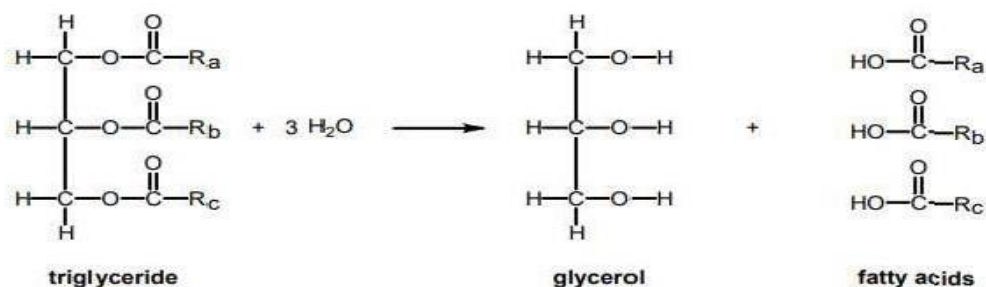
Triglycerides can be broken apart under basic conditions (a saponification reaction) to produce long-chain carboxylate salts



Example:



Triglycerides (fats) can be hydrolyzed to produce glycerol and 3 fatty acids in the presence of acid and heat or with a suitable lipase enzyme under biological conditions.



This hydrolysis reaction produces glycerol and fatty acids, which are carboxylic acids derived from fats and oils. In the fatty acids, R_a , R_b , and R_c , represent groups of carbon and hydrogen atoms in which the carbon atoms are attached to each other in an unbranched chain.

The hydrolysis reaction is promoted by acids and by bases. When a strong base such as NaOH (lye) is used, the product contains salts of the fatty acids. These salts of fatty acids are the functional ingredient in soap. The ingredients lists of some soaps include sodium tallowate, a generic name for the mixture of fatty acid salts obtained from tallow (animal fat), and sodium cocoate, obtained from coconut oil.

Hardening of oils

Chemically, the hydrogenation of oils is the reduction of the double bonds in unsaturated fatty acids to single saturated bonds, by the reaction of hydrogen gas in the presence of a metal catalyst.

Rancidity

Rancidity is a condition in which fat attains a bad taste and disagreeable odour.

Types of rancidity

There are 2 types of rancidity:

1. Hydrolytic rancidity
2. Oxidative rancidity

Hydrolytic rancidity: Fats are hydrolyzed into glycerol and fatty acids in presence of moisture, warm temperature and bacterial enzymes.

Oxidative rancidity: It occurs by oxidation of unsaturated fatty acids present in fats and oils forming lipid peroxides, fatty aldehydes, ketones and short chain fatty acids.

Rancidity is predisposed by: light, moisture, warm and temperature

Rancidity leads to:

- ✚ Fats and oils attain bad taste.
- ✚ Fats and oils attain disagreeable odour.
- ✚ Production of toxic compounds as lipid peroxides, aldehydes and ketones.

Prevention of rancidity:

- Avoid exposure of fats to light, moisture and high temperature.
- Addition of antioxidants to fats and oils.

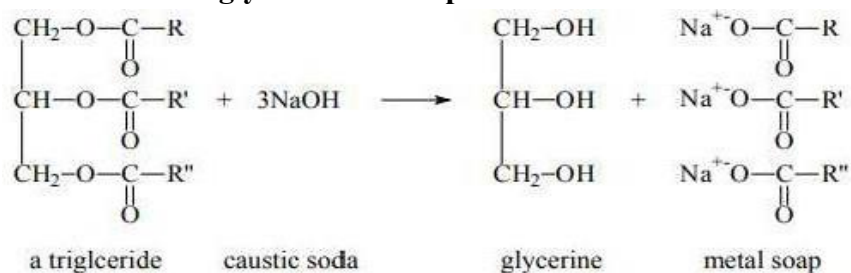
6.1.4. Soaps and detergents

Soaps are water-soluble sodium or potassium salts of fatty acids. Soaps are made from fats and oils, or their fatty acids, by treating them chemically with a strong alkali.

Saponification of fats and oils is the most widely used soapmaking process.

Soaps are the product of the reaction between a fat and sodium hydroxide:

fat + 3NaOH → glycerine + 3 soap

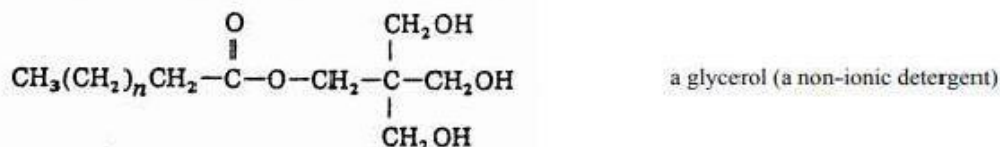
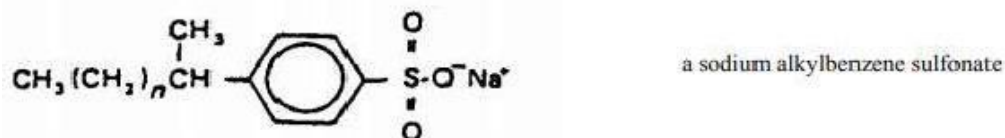


Detergent

Detergents are complex formulations that contain more than 25 different ingredients, which can be categorized into the following main groups:

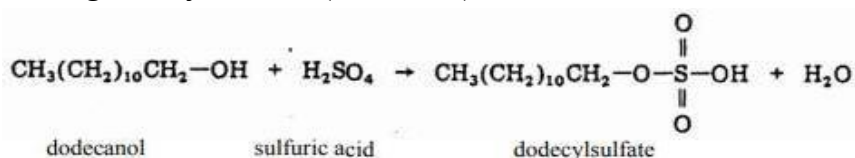
- Surfactants
- Bleaching agents
- Builders
- Additives

Detergents are structurally similar to soaps, but differ in the water-soluble portion. Three examples of detergents are shown below.



Preparation of a detergent

A synthetic detergent, a sodium alkyl sulfate called sodium dodecylsulfate, will be prepared by reacting dodecyl alcohol (dodecanol) with sulfuric acid.



When a soap or detergent is added to water, a polar solvent, the molecules form clusters, known as **micelles**, in which the polar ends of the molecules are on the outside of the cluster and the non-polar ends are in the middle.

The cleaning action of both soaps and detergents results from their ability to emulsify or disperse water-insoluble materials (dirt, oil, grease, etc.) and hold them in suspension in water. This ability comes from the molecular structure of soaps and detergents. When a soap or detergent is added to water that contains oil or other water-insoluble materials, the soap or detergent molecules surround the oil droplets. The oil or grease is “dissolved” in the alkyl groups of the soap molecules while the ionic end allows the micelle to dissolve in water. As a result, the oil droplets are dispersed throughout the water (this is referred to as emulsification) and can be rinsed away.

6.1.5. Comparison of Organic Compounds

Property	Alkane	Alkene	Alkyne	Benzene	Carboxylic Acid	Ester	Aldehyde	Ketone
Common Name	Paraffin	Olefin	Acetylene series	Benzol	Formic acid (HCOOH)	Ethyl acetate (EtOAc)	Formaldehyde (HCHO)	Acetone CH_3COCH_3
IUPAC Name	Methane, ethane, etc.	Ethene, Propene, etc.	Ethyne, Propyne, etc.	Benzene	Methanoic acid, etc.	Ethyl ethanoate, etc.	Methanal, Ethanal, etc.	Propanone, etc.
General formula	$\text{C}_n\text{H}_{2n+2}$	C_nH_{2n}	$\text{C}_n\text{H}_{2n-2}$	C_6H_6	R-COOH	R-COO-R	R-CHO	R-CO-R'
Classification	Saturated hydrocarbons	Unsaturated hydrocarbons	Unsaturated Hydrocarbon	Aromatic hydrocarbons	Organic acids	Organic esters	Organic carbonyls	Organic carbonyls
Physical Properties	Nonpolar, low MP & BP, insoluble in water	Nonpolar, low MP & BP, slightly reactive	Nonpolar, linear, higher BP than alkenes	Aromatic, nonpolar, volatile	Weak acid, high BP, soluble	Sweet smelling, volatile, nonpolar	Pungent smell, polar	Moderate polarity, pleasant smell
Chemical Properties	Combustion, Substitution	Addition, Combustion	Addition, Combustion	Substitution	Acidic, forms esters	Hydrolysis, Reduction	Oxidation, Reduction	Oxidation, Reduction
Uses	Fuels, lubricants	Plastic synthesis	Welding, fuel gas	Solvent, fuel, drugs	Food, medicine	Perfumes, flavors	Preservatives, disinfectants	Solvents, pharmaceuticals
Preparation	Crude oil refining	Cracking of alkanes	Dehydro-halogenation	Coal tar, catalytic reforming	Oxidation of alcohols	Esterification of acids & alcohols	Oxidation of alcohols	Oxidation Of alcohols

6.2. Questions with explanation

256. In hydrocarbons (non-polar compounds), the one with the lowest molar mass has the lowest melting and boiling point; while the one with the highest molar mass has the highest boiling point. Which of the following is a correct order of decreasing (from highest to lowest) melting and boiling point?

- A. $\text{CH}_4 < \text{C}_2\text{H}_6 < \text{C}_3\text{H}_8 < \text{C}_4\text{H}_{10} < \text{C}_5\text{H}_{12}$
B. $\text{C}_5\text{H}_{12} < \text{CH}_4 < \text{C}_2\text{H}_6 < \text{C}_4\text{H}_{10} < \text{C}_3\text{H}_8$

- C. $\text{CH}_4 > \text{C}_2\text{H}_6 > \text{C}_3\text{H}_8 > \text{C}_4\text{H}_{10} > \text{C}_5\text{H}_{12}$
D. $\text{C}_5\text{H}_{12} > \text{C}_4\text{H}_{10} > \text{C}_3\text{H}_8 > \text{C}_2\text{H}_6 > \text{CH}_4$

Brief:

- The stronger the intermolecular attractive forces, the more energy that is required to break them apart.
 - Larger molecules have higher boiling and melting points.
 - In general, larger molecules have higher boiling points than smaller molecules of the same kind, indicating that dispersion forces increase with mass, number of electrons, number of atoms or some combination thereof.
- The boiling point of organic molecules is in the order of:
- Alkanes < Alkenes < Alkynes < Haloalkanes < Ether < Aldehyde < Ketone < ester < Alcohol < Carboxylic Acid.

257. Detergents are preferable to hard soap for washing (cleaning process). This because detergents

- A. Made with chemicals that are not affected by hard water.
- B. Work equally well with hard and soft water
- C. Are cheaper in price than hard soap
- D. Are manufactured more easily than soap

✚ Detergents are preferred over soap due to the following reasons:

- Detergents act more effectively on hard water as compared to soaps
- The sulfonate group does not attach itself to the ions present in hard water.
- They do not form insoluble precipitates with the dissolved calcium and magnesium ions in hard water
- Detergents have strong cleansing properties.
- The formation of more lather is observed in hard water than in soft water in the case of detergents.
- They do not lose their effectiveness in hard water and saline water.

258. Which of the following is found in all alcoholic beverages and used as organic solvents?

- A. $\text{C}_3\text{H}_7\text{OH}$ B. CH_3OH C. $\text{C}_2\text{H}_5\text{OH}$ D. $\text{C}_5\text{H}_{11}\text{OH}$

Brief:

✚ Ethanol ($\text{C}_2\text{H}_5\text{OH}$) is the 2nd member of alcohol that is the major constituent of any alcoholic beverage with different percentage.

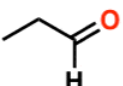
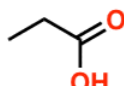
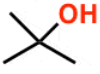
259. Alcohols that yield Ketone on oxidation are classified as

- A. dihydric alcohols C. secondary alcohols
B. trihydric alcohols D. primary alcohols

Brief

✚ Primary alcohol gets easily oxidized to an aldehyde and can further be oxidized to carboxylic acids ✚ Secondary alcohol gets easily oxidized to ketone but further oxidation is not possible.

✚ Tertiary alcohol doesn't get oxidized in the presence of sodium dichromate.

The bottom line		
Starting material	"Weak" oxidant <i>PCC, DMP, Swern, etc</i>	"Strong" oxidant <i>KMnO₄, H₂CrO₄ (or equivalent)</i>
 Primary alcohol	 Aldehyde	 Carboxylic acid
 Secondary alcohol	 Ketone	 Ketone
 Tertiary alcohol	No reaction	No reaction

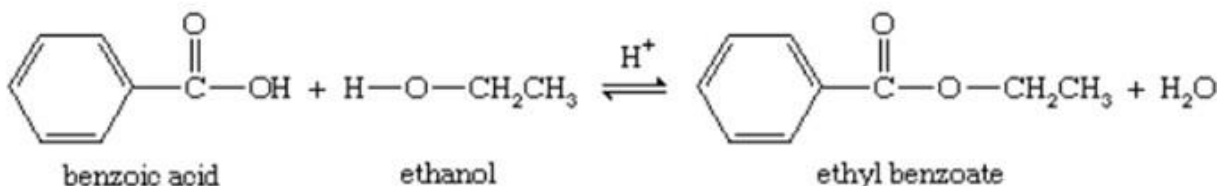
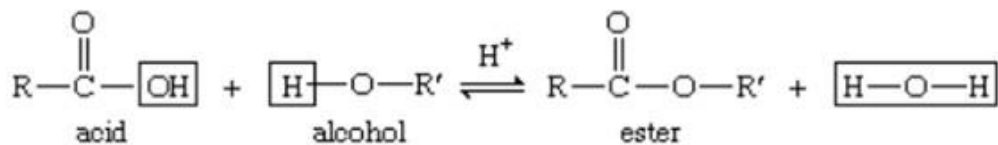
260. Which of the following would react to form ethyl ethanoate?

- A. Ethanol and Ethanoic acid C. Ethanol and Butanoic acid
B. Methanol and Ethanoic acid D. Ethanol and Propanoic acid

Brief:

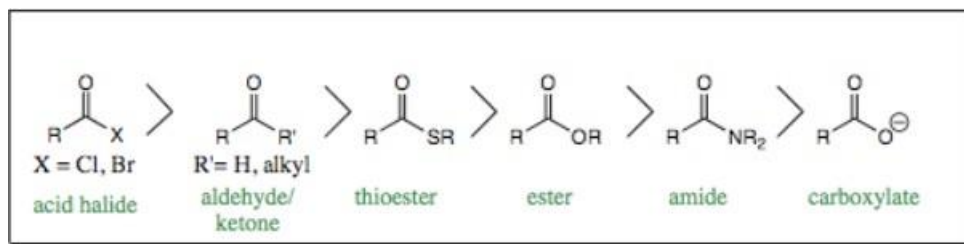
✚ The reaction, called **Fischer esterification**, is characterized by the combining of an alcohol and an acid (with acid catalysis) to yield an ester plus water.

✚ Under appropriate conditions, inorganic acids also react with alcohols to form esters.

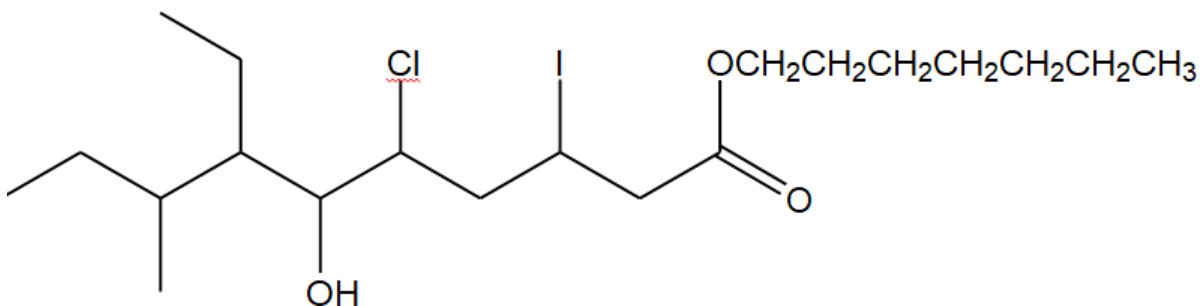


261. Which of the following statement is true? (self-practice)
- A. Aldehydes are oxygen containing hydrocarbon those are obtained when secondary alcohols are oxidized by $\text{K}_2\text{Cr}_2\text{O}_7$ in basic condition
 - B. Ketones are oxygen containing hydrocarbon those are obtained when primary alcohols are oxidized by $\text{K}_2\text{Cr}_2\text{O}_7$ in basic condition
 - C. Carboxylic acids are oxygen containing hydrocarbon those are obtained when primary alcohols are oxidized by $\text{K}_2\text{Cr}_2\text{O}_7$ in acidic condition
 - D. Carboxylic acids are oxygen containing hydrocarbon those are named by replacing the last terminal -e of the corresponding alkanes in to -al
262. Which of the following statement is correct?
- A. Esters are oxygen containing hydrocarbons those are highly reactive than aldehydes
 - B. Ketones are oxygen containing hydrocarbons those are less reactive than esters
 - C. Esters are oxygen containing hydrocarbons those are stabilized by resonance
 - D. Carboxylic acids are oxygen containing hydrocarbons those have no hydrogen bonding

Brief:



263. What is the correct IUPAC name for the molecule given below? (self-practice).



- A. hexyl-5-chloro-6-hydroxy-7-ethyl-3-iodo-8-methyldodecanoate
- B. heptyl-5-chloro-7-ethyl-6-hydroxy-3-iodo-8-methyldecanoate..
- C. heptyl-5-chloro-7-methyl-6-hydroxy-3-iodo-8-ethyldecanoate.
- D. hexyl-7-methyl-5-chloro-m-iodine-6-aqua-8-ethyldecanoate.

264. Which trend in increasing order of the boiling point of the following acids is correctly stated?

* $\text{CH}_3\text{CH}_2\text{COOH}$, $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}$, $\text{CH}_3\text{CH}(\text{CH}_3)\text{CH}_2\text{COOH}$ and $\text{CH}_3(\text{CH}_2)_5\text{COOH}$.

- A. $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH} < \text{CH}_3(\text{CH}_2)_5\text{COOH} < \text{CH}_3\text{CH}_2\text{COOH} < \text{CH}_3\text{CH}(\text{CH}_3)\text{CH}_2\text{COOH}$
- B. $\text{CH}_3(\text{CH}_2)_5\text{COOH} < \text{CH}_3\text{CH}_2\text{COOH} < \text{CH}_3\text{CH}(\text{CH}_3)\text{CH}_2\text{COOH} < \text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}$
- C. $\text{CH}_3\text{CH}_2\text{COOH} < \text{CH}_3\text{CH}(\text{CH}_3)\text{CH}_2\text{COOH} < \text{CH}_3\text{CH}_2\text{CH}_2\text{COOH} < \text{CH}_3(\text{CH}_2)_5\text{COOH}$
- D. $\text{CH}_3\text{CH}_2\text{COOH} < \text{CH}_3\text{CH}_2\text{CH}_2\text{COOH} < \text{CH}_3\text{CH}(\text{CH}_3)\text{CH}_2\text{COOH} < \text{CH}_3(\text{CH}_2)_5\text{COOH}$

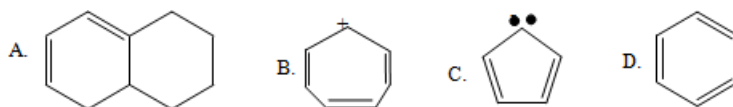
Brief:

Alcohols have the highest boiling point because of hydrogen bonding. Carbonyl compounds have dipole dipole interaction which means the boiling point is high, simply not as high as that of an alcohol.

➤ The boiling point of organic molecules increases in the order of

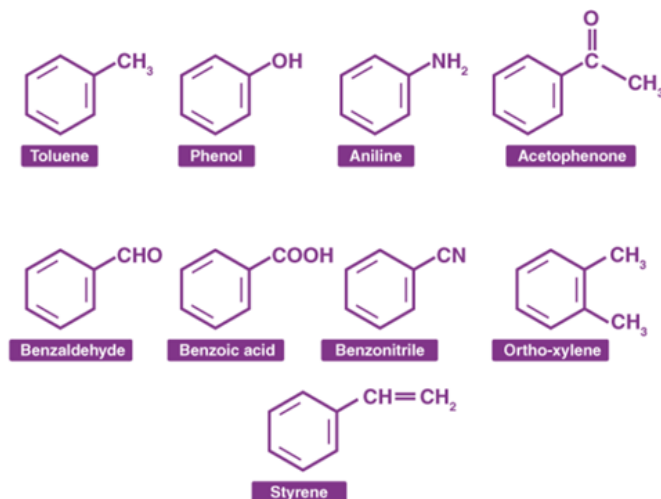
✚ Alkanes < Alkenes < Alkynes < Haloalkanes < Ether < Ester < Amine < Aldehyde
Ketone < Alcohol < Carboxylic Acid

265. Which of the following molecule is aromatic?



Brief:

❖ Aromatic hydrocarbon, are hydrocarbons containing sigma bonds and delocalized pi electrons between carbon atoms in a ring. For example, benzene. They are known as aromatic due to their pleasant smell.



- ❖ Aromatic compounds are broadly divided into two categories: benzenoids (one containing benzene ring) and non-benzenoids (those not containing a benzene ring) for example, furan. Any hydrocarbon can be classified as an aromatic compound provided they follow the Huckel rule.
- ❖ According to Huckel rule, for a ring to be aromatic it should have the following properties:
 1. Planarity structure
 2. Complete delocalization of the π electrons in the ring
 3. Presence of $(4n + 2)$ π electrons in the ring where n is an integer ($n = 0, 1, 2, \dots$)

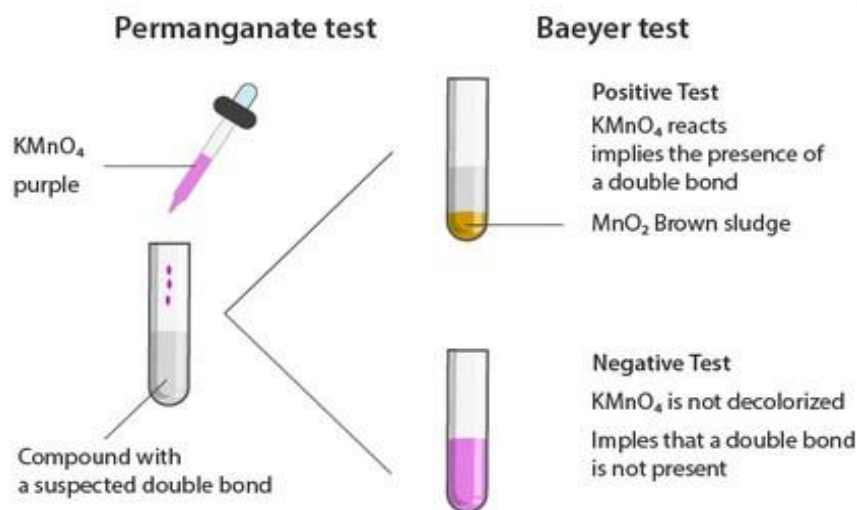
266. Sifen is a grade 10 student. She took three different test tubes and added 5.00ml of toluene (methylbenzene) to each test tube. Then she added 2 drops of Br_2 water in the first test tube, equal drops of KMnO_4 in the second test tube and again equal drops of concentrated H_2SO_4 in the third test tube. In which test does a reaction occur?

- A. Test tube 1 and 2 B. Test tube 2 and 3 C. test tube 1,2 and 3. D.test tube 1 and 3

Brief:

- There are two methods for detecting the unsaturation in an organic compound.
 1. Alkaline potassium permanganate test (Baeyer's test)

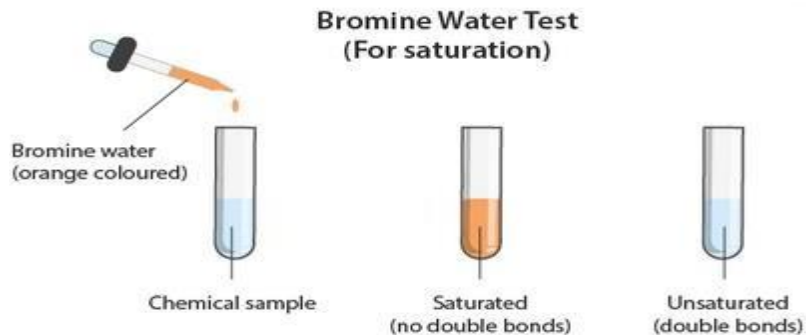
The pink colour potassium permanganate disappears when an alkaline potassium permanganate is added to an unsaturated hydrocarbon. The disappearance of pink colour may take place with or without the formation of brown precipitate of manganese oxide. *So, decolourization of pink colour of KMnO_4 indicates unsaturation.*



2. Bromine test

Bromine solution is brown. In this test when bromine solution is added to the unsaturated hydrocarbon the brown colour disappears if the hydrocarbon is unsaturated. Bromine forms an addition product with the unsaturated hydrocarbon.

So, decolourization of orange-red colour of bromine solution indicates unsaturation.



NB:

- ✓ Baeyer's reagent is a cold potassium permanganate alkaline solution, which is a powerful oxidant that makes it a redox reaction
- ✓ Aldehydes decolourise potassium permanganate solution due to its oxidation to carboxylic acid. Bromine water test can be carried out in this case to find out unsaturation.

- ✓ The double bond breaks when alkenes react with bromine, and each carbon is attached to a bromine atom. It gives a colourless liquid, the bromine loses its original red-brown colour.

267. Natural gas is mainly composed of:

- A. Gases such as nitrogen and hydrogen sulphide
- C. Gases such as hydrogen and sulphur
- B. Methane gas
- D. All oxygen containing hydrocarbons

Brief:

- Natural gas is an odorless, gaseous mixture of hydrocarbons—predominantly made up of methane (CH₄).
- Methane is also known as the Marsh Gas for which Natural Gas (CNG) consists mainly of methane (95%), a relatively unreactive hydrocarbon and makes its nearly complete combustion possible. It is a greenhouse gas.

268. Which of the following set of organic compounds are all organic

- A. Petroleum, benzene, paraffin, paracetamol, fertilizer
- B. TNT, Toluene, Naphthalene, Olefin, benzene
- C. Benzene, Petroleum ether, Gastric juice, amino acid
- D. Citric acid, table salt, Oxalic acid, butyric acid, sedative

Brief:

- TNT (Trinitrotoluene) is an organic compound.
- Toluene is an aromatic hydrocarbon.
- Naphthalene is an aromatic hydrocarbon.
- Olefin (another name for alkene) is an unsaturated hydrocarbon.
- Benzene is an aromatic hydrocarbon.
- Petroleum, benzene, and paraffin, and paracetamol, fertilizer: Not all of these are purely organic. Petroleum is a mixture of hydrocarbons, and fertilizer might contain inorganic compounds.
- Benzene, Petroleum ether, Gastric juice, amino acid: Gastric juice is not organic; it contains mainly water and hydrochloric acid, both of which are inorganic.
- Citric acid, table salt, Oxalic acid, and butyric acid, sedative: Table salt (sodium chloride) is an inorganic compound, though the rest are organic.

269. The relative spatial arrangement of atoms or groups of atoms about the double bond is known as:

- A. Conformational isomerism C. Position isomerism
B. Linkage isomerism D. Geometrical isomerism

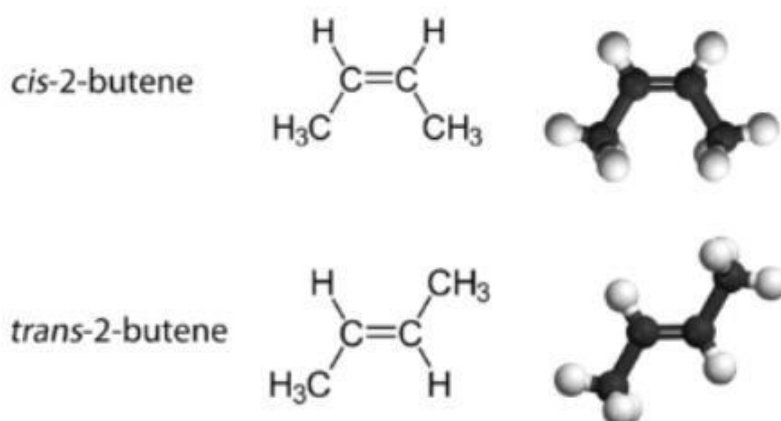
Brief:

- ✓ Alkenes show both structural isomerism and geometrical isomerism.

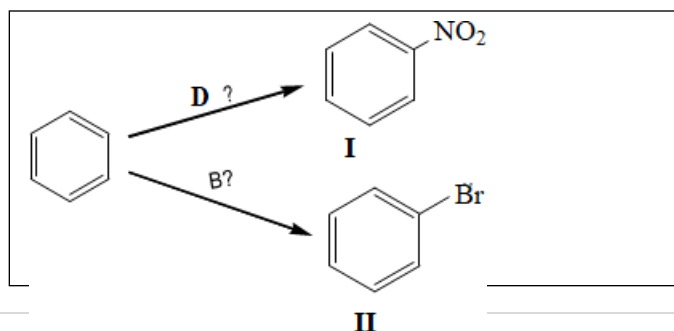
A structural isomer also known as a constitutional isomer is one in which two or more organic compounds have the same molecular formulas, but different structures..

Structural isomerism is applicable for both alkane and alkene

- ✓ The two common structural isomerism of alkenes are position and chain isomerism
✓ Position isomers have same molecular formula, the same location of substituents but different possible position of double bond.
✓ Chain isomerism; have the same position of double bond but different location of substituents
✓ Geometric isomers are isomers in which the order of atom bonding is the same, but the arrangement of atoms in space is different. The double bond in an alkene is not free to rotate because of the nature of the pi bond. Therefore, there are two different ways to construct the 2-butene molecule



270. In laboratory Samond carried out an experiment on substitution reaction starting from C_6H_5 . He obtained $C_6H_5NO_2$ and C_6H_5Br as shown in the reaction.



What are the missed reagents represented by letters **D** and **B** respectively.

- A. HNO_3 , H_2CO_3 & Br_2
- B. H_2 , Ni, heat & Br_2
- C. HBr , and CH_3OH
- D. HNO_3 , H_2SO_4 , heat & Br_2 , AlBr_3

Brief:

Benzene reacts with chlorine or bromine in substitution reaction, but only in the presence of a catalyst. The catalyst is either aluminum chloride (or aluminum bromide if you are reacting benzene with bromine) or iron.

Halogenation:	C_6H_6	+ Cl_2 & heat	FeCl_3 catalyst
Nitration:	C_6H_6	+ HNO_3 & heat	H_2SO_4 catalyst
Sulfonation:	C_6H_6	+ H_2SO_4 + SO_3 & heat	
Alkylation:	C_6H_6	+ R-Cl & heat	AlCl_3 catalyst

UNIT SEVEN

7. ACID BASE EQUILIBRIA

Unit outcome

➤ ***At the end of this unit, you will be able to:***

- ❖ *Define acid and bases based on Arrhenius definition, Bronsted lowery definition, Lewis definition.*
- ❖ *Describe the dissociation of water, weak mono-protic and polyprotic acids, and weak bases*
- ❖ *solve equilibrium problems involving concentration of reactants and products, a , K_b , P_H and*
- ❖ *Explain the common ion effect, buffer solution, hydrolysis of salts, acid-base indicators and acid-base titrations*
- ❖ *Calculate p_H , pOH , K_a , K_b and other quantities of acid base equilibrium.*
- ❖ *Determine the equivalents of acid or base that are required to neutralize specific amount of acid or base*
- ❖ *explain how to solve problems involving concentration and p_H of acid base titration*
- ❖ *Identify suitable indicators for different types of titration*

7.1. Introduction

Acids and bases are the most prominent inorganic compounds. Some of the most important processes in chemical and biological systems are acid base concepts. Since they are too familiar in real life situations, used to many industrial processes. A number of familiar materials, such as aspirin and vitamin C are acidic. Vinegar is a dilute solution of acetic acid, lemonade contains citric acid, the battery in the car contains sulphuric acid. Milk of magnesia is basic, as are baking soda and washing soda.

Generally, acids and bases are the most important classes of compounds and necessary in biology and chemistry environment.

- Acids & Bases Found in Homes
- ✓ Citrus Fruits- Oranges and lemons, for instance, contain citric acid, which makes them acid home products.
- ✓ Toothpaste- Toothpaste is a Base, and it has an alkaline composition. As soon as we eat, the food breaks down and releases acid.
- ✓ Carbonated Beverages. –carbonic Acid
- ✓ Vinegar – weak acetic acid
- ✓ Lemon juice – citric acid and some ascorbic acid
- ✓ Any citrus fruit – citric acid and some ascorbic acid
- ✓ Most other fruits – citric acid, possibly tartaric, oxalic, or malic acid
- ✓ Carbonated soda – phosphoric, carbonic, and sometimes citric acid
- ✓ Battery acid – sulfuric acid
- ✓ Aspirin – acetylsalicylic acid
- ✓ Muriatic acid – hydrochloric acid
- ✓ Sour candies – citric acid
- ✓ Yogurt – lactic acid
- ✓ Brewed coffee or tea- Tannic acid
- ✓ Most alcoholic beverages- Acetic acid
- ✓ Stomach acid and vomit – hydrochloric acid
- ✓ Urine (only slightly acidic)
- ✓ Baking soda – sodium bicarbonate
- ✓ Washing soda – sodium carbonate
- ✓ Soap – Either sodium or potassium hydroxide
- ✓ Chalk – calcium carbonate

7.1.1. Definition of Acids and Bases

The term 'acid' is derived from the Latin word 'ACERE' which means sour while base is derived from a Latin word "alquili", which means ash of plant.

Traditional definition of Acid and base

An acid is any substance that in water solution tastes sour, changes blue litmus paper to red, reacts with some metals to liberate hydrogen, reacts with bases to form salts, and promotes chemical reactions (acid catalysis). The basic definition of Acids and Bases are categorized to three classes. Arrhenius definition, Bronsted lowry definition and Lewis definitions will study next.

1. Arrhenius Definitions

The Swedish chemist **svante Arrhenius** framed the first successful concept of Acids and Bases. The Arrhenius definition states that **an acid produces H^+ in solution and a base produces OH^-** . This theory was developed by Svante Arrhenius in 1883. Later, two more sophisticated and general theories were proposed. These are the Brønsted-Lowry and the Lewis definitions of acids and bases.

According to Arrhenius,

An acid is a substance that, when dissociate in water, increase the concentration of hydroium ion, $H_3O^+(aq)$

A base is a substance that, when dissociate in water increase the concentration of hydroxide ion $OH^-(aq)$

For more general, Arrhenius concept is aqueous dependent, some typical Arrhenius Acids are HCl , HNO_3 , H_2SO_4 , $HClO_4$ and some typical bases are NH_3 , $NaOH$, $Ca(OH)_2$, $Ba(OH)_2$,

$HCl(aq) + H_2O(l) \rightarrow Cl^-(aq) + H^+(aq)$: HCl is Arrhenius

Acid $NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$: NH_3 is

Arrhenius base

Despite in success, the Arrhenius definition of acid and base is limited to *aqueous* solutions. but there are number of reactions that occurs in other solvents and in the gas phase, example, when benzene solution of Acetic Acid and ammonia are mixed, acetic acid react with ammonia to form ammonium acetate,

$CH_3COOH(l) + NH_3(g) \rightarrow CH_3COONH_4(s)$

Ammonia and hydrogen chloride react as a base and an acid in the gas phase to form ammonium chloride.

$NH_3(g) + HCl \rightarrow NH_4Cl(s)$

Therefore, a more general definition of acids and bases that include reaction in the gas phase and in nonaqueous solvent is needed.

Advantages of Arrhenius Theory

The various advantages of Arrhenius theory are that it provides knowledge about:

- It was also the first theory on acids and bases that acknowledged the significance of solvents inneutralization and water formation during the reaction.
- The properties of acids and bases
- The strength of acids and bases
- Concept of neutralization and hydrolysis

Limitations of Arrhenius Theory

1. Arrhenius Theory is limited to the study of acids and bases in aqueous solution only and notapplicable in gaseous and non-aqueous solutions.
2. Arrhenius theory is applicable only to acids having ionizable hydrogen (HA). For example: HCl , HBr . It is not applicable for SO_2 , CO_2 , and $AlCl_3$.
3. Arrhenius theory is only applicable to bases having the formula BOH . For example: $NaOH$, KOH . It is not applicable for Na_2CO_3 , NH_3 and pyridine. In fact, the need to prove the relationship between base and OH^- ion led to the proposal of the formula NH_4OH as base of ammonia in water, the actual formula being NH_3 .
4. Arrhenius definition is failed to explain the extent of definition
5. This theory fails to explain the properties of acids and bases in any solvents)like

benzene or acetone) or gases, other than water.

6. According to Arrhenius, all salts should produce solutions that are neither acidic nor basic.

However, there are some exceptions to this and Arrhenius theory fails to explain this, e.g

- ✓ Equal amounts of HCl and ammonia yield a slightly acidic solution
 - ✓ Equal amounts of acetic acid and sodium hydroxide yield a slightly basic solution
7. According to this idea, solvents play no role in determining the acidic or basic constitution of substances. For instance, hydrochloric acid is weak when dissolved in benzene but strong when dissolved in water.
 8. According to this theory, hydrogen ions release in the aqueous solution and remain there, increasing the concentration of hydrogen ions in the solution. However, this is not true because hydrogen can never exist freely as hydrogen ions. It exists as hydronium ion (H_3O^+).
(i.e. failed to explain the existence of hydrogen ion)

2. Bronsted Lowry Definition

A somewhat more general approach was that proposed independently in 1923 by the Danish chemist,

J.N. Bronsted and the British chemist *T.M. Lowry* according to Bronsted Lowry definition;

- An acid is proton donor, any substance that can donate a proton, H^+ is an acid which must contain H in its formula; HCl, H_2SO_4 , HF, HNO_3 , HClO_4 are examples of acids.
- All Arrhenius acids are Bronsted Lowry acids, but the reverse may not be true
- A base is proton acceptor, any species that can accept a proton is base.
- A base must contain a lone pair of electrons to bind the H^+ , eg NH_3 , F^- , HSO_4^- , HCO_3^- , H_2PO_4^- , HPO_4^{2-}

Example $\text{H}_2\text{S} + \text{NH}_3 \rightarrow \text{HS}^- + \text{NH}_4^+$;

In the forward reaction, H_2S act as acid by donating proton, H^+ to NH_3 , which act as a base

- ❖ In the reverse reaction, NH_4^+ act as acid while HS^- act as a base. Notice that, in forward reaction, the acid H_2S , become a base, HS^- and the base NH_3 , become an acid NH_4^+ . According to Bronsted terminology, H_2S and HS^- are conjugated acid-base pair. i.e. HS^- is the conjugated base of acid, H_2S . Hence, when acid loses a proton, form a conjugated base. similarly, NH_3 and NH_4^+ are a conjugated acid - base pairs. NH_4^+ is the conjugated acid of the base NH_3 , hence, when a base accept a proton, it forms a conjugated acid of base. A conjugate acid, base pairs consist of two species that differ from each other by the presence of one H^+ unit. i.e.
 - the conjugated base of the pair has one fewer H and one more negative charge than the acid
 - The conjugated acid of the pair has one more H and one fewer negative charge than the base.

A *Bronsted-lowry* acid-base reaction occurs when an acid and a base react to form their conjugate base and conjugate acid respectively;



- ✓ Base1 is the conjugated base of Acid1
- ✓ acid 2 is the conjugated acid of base2

Examples: $\text{HCOOH(l)} + \text{CN}^{\text{-(aq)}} \rightleftharpoons \text{HCOO}^{\text{-(aq)}} + \text{HCN(l)}$

Acid	Base	Conjugated acid	Conjugated base	Conjugated Acid-Base pair
HCOOH	CN ⁻	HCN	HCOO ⁻	HCOOH/HCOO ⁻ , CN ⁻ /HCN

Limitations of Brønsted–Lowry Theory

- ❖ Brønsted–Lowry Theory concept cannot explain the reactions occurring in non-protic solvents such as COCl_2 , SO_2 , N_2O_4 , etc.
- ❖ It cannot explain reactions of acid and basic oxides which can take place even in the absence of solvents.
- ❖ It does not explain the reactions between acidic oxides and basic oxides. $\text{CaO} + \text{CO}_2 \rightarrow \text{CaCO}_3$

3. Lewis concept of acid and base

➤ Accordingly:

An Acid is a species with vacant orbital capable of accepting an electron pair

A Base is a species with an excess of electron which can donate an electron pair to form coordinative covalent bond.

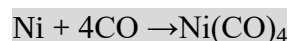
Example: $\text{BF}_3 + \text{NH}_3 \rightarrow \text{BF}_3\text{NH}_3$; BF_3 is an acid (accept electron) while NH_3 is base (donate electron)

A proton itself is *lewis* acid. Lewis concept is the most general one/

NB; Substance that are classified as *Bronsted Lowry* base are also bases according to Lewis system. Chemical substance that can include the following Lewis concepts

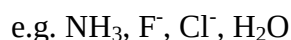
- I. Molecules or atoms that have vacant orbital Or an incomplete octet are acid e.g. BF_3 , AlCl_3
- II. Many simple cat ions are Lewis Acids. Example, Fe^{3+} has vacant orbital and reacts strongly with cyanide ion to form the ferric cyanide ion, $\text{Fe}^{3+} + 6\text{CN}^- \rightarrow [\text{Fe}(\text{CN})_6]^{3-}$. Thus Fe^{3+} has a vacant orbital that accept pair of electrons donated by CN^-
- III. Some compounds with multiple bonds can behave as lewis acid

IV. Some metal can function as acids in the formation of compound such as carbonyls, which are produced by the reaction of metal with carbon monoxide is lewis acid :



V. Compounds that have central atoms capable of expanding their valance shells are Lewis Acid e.g. SiCl_4 , PF_5 , SbF_5

Any species which is negatively charged or having lone pair are Lewis base



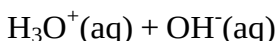
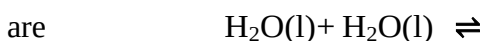
The following table will summarize the three definitions of acid and base

Types of	Definition	Bases of definition	Description
Arrhenius definition		Aqueous dependent	<i>Acid</i> increase concentrations of H^+ <i>Base</i> increase concentration of OH^-
Bronsted Lowry definition		proton transfer	<i>Acid</i> are proton donor <i>Base</i> are proton acceptor
Lewis definition		Electron transfer	<i>Acids</i> are e- pair acceptor <i>Bases</i> are e- pair donor

Auto-ionization and Ion-product of water

❖ ***Auto-ionization*** is the reaction process occurs between two the same species one act as An acid while other act as a base.

The familiar example for such rxns



For the reaction ; $\text{H}_2\text{O}(\text{l}) + \text{H}_2\text{O}(\text{l}) : \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$, according to Bronsted lowry theory, the reaction is viewed as a proton transfer from one water molecule to other. Since water act as both acid & base called amphoteric. The reaction is reversible, and H_2O , H_3O^+ and OH^- are at equilibrium. In pure water at 25°C ,

$[\text{H}^+] = [\text{OH}^-] = 1 \times 10^{-7}$. These are low concentration and tell us that very few H_2O molecule are ionized and that the equilibrium lie to the left at equilibrium. the product of $[\text{H}^+]$ and $[\text{OH}^-]$ equals constant called ion product constant for water

- ❖ At 25°C , $K_w = 1 \times 10^{-14}$ and $[\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14}$ for dissociation of pure water.
- ❖ The value of K_w applies to all aqueous solution at 25°C when an acid is added to water, the $[\text{H}^+]$ increase.
- ❖ Therefore the $[\text{OH}^-]$ must decrease in order for K_w to remain constant.
i.e in acidic solution, $[\text{H}^+] > [\text{OH}^-]$, similarly when a base is added to water and $[\text{OH}^-]$ increase the $[\text{H}^+]$ must decrease for the same reason.

Acidic, basic and neutral solutions are characterized as follows

<u>Neutral solution</u>	<u>Acidic Solution</u>	<u>Basic Solution</u>
$[H^+] = [OH^-]$	$[H^+] > [OH^-]$	$[OH^-] > [H^+]$
$[H^+] = 1 \times 10^{-7}$	$[OH^-] < [H^+]$	$[OH^-] > 1 \times 10^{-7}$
$[OH^-] = 1 \times 10^{-7}$	$PH < 7$	$[H^+] < 1 \times 10^{-7}$
$PH=7$	$H^+] > 1 \times 10^{-7}$	$POH < 7$
$POH=7$	$PH < 7$	$PH > 7$
$PH=POH$	$POH > 7$	$PH > POH$
	$PH < POH$	

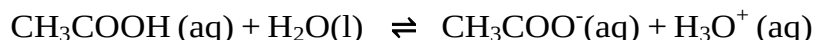
- ❖ For two different acids of equal concentrations, the stronger acids produce, the greater quantity of hydrogen ion. Thus the concentration of hydrogen ion is a measure of relative strength of two acid similarly, stronger base will contain a large concentration of hydroxide ion than the solution of weaker base if both solution are of equal concentrations.

7.1.2. Common ion effect and buffer solution

Common ion effect

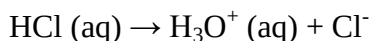
- Common ion effect is shifting of ionic equilibrium caused by the addition of absolute that provides an ions that takes part in equilibrium
- Consider the solution of acetic acid, CH_3COOH
- $CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$

- ❖ Suppose we add strong electrolyte substance which contains CH_3COO^- , CH_3COONa



$CH_3COONa \rightarrow CH_3COO^-(aq) + Na^+$ on addition of CH_3COONa , the concentration of acetate ion increase as a result it consumes more H_3O^+ . According to **Lechatlier** principle the equilibrium lies to the left as a result dissociation extent of Acetic acid decrease. This suppression of the dissociation of acetic acid is called *common ion effect*.

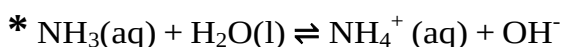
For the dissociation of formic acid $HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+$ what is the effect of adding strong acid, HCl on dissociation reaction



* H_3O^+ is common ion So on addition of HCl

- $[H_3O^+]$ increase
- PH decrease
- Acidity increase

* What is the effect of adding NH_4Cl to dissociation of NH_3 ?



* $NH_4Cl(aq) \rightarrow NH_4^+ + Cl^-$, So on addition of NH_4Cl

- $[\text{NH}_4^+]$ increased
- $[\text{OH}^-]$ decreased, more OH^- decreased more OH^- consumed
- Equilibrium shift to left
- POH increase
- PH decrease
- Acidity increase
- Basicity decrease

Buffer solution

- Any solution like the preceding one, which contains both weak acid (HF) and its conjugate base (F^-) has ability to neutralize added strong acid and bases with very little change in PH such solution is called *buffer solution*.
- Because it resists significant changes in the PH value a buffer must contains an acid to react with any OH^- ions that may be added and a base to react with any added H^+ ions. Buffer solution is prepared from weak Acid and its conjugated base and weak base and its conjugate acid
- A buffered solution is one that resists a change in its pH when either hydroxide ions or protons (H_3O^+) are added. Very little change in pH is witnessed even when a strong acid or base is added.
- Buffered solutions are solutions of weak acids or bases containing common ion.
- When a strong acid or base is added to a buffered solution, do the stoichiometric calculation first.
- Buffer solution is prepared from weak acid and its salt or weak bas with its salt

e.g. $\text{NH}_3/\text{NH}_4\text{Cl}$
 $\text{CH}_3\text{COOH}/\text{CH}_3\text{COONa}$
 HF/NaF

When a small amount of strong acid is added to buffer, it is neutralized by the weak base. Also when a small amount of strong base is added to buffer it is neutralized by the weak base.

E.g. In HA/F^- buffer

- On addition of strong base, the neutralization reaction takes place as follows
- $\text{OH}^- + \text{HF} \rightarrow \text{H}_2\text{O} + \text{F}^-$
- On addition of strong acid the neutralization reaction is $\text{H}^+ + \text{F}^- \rightarrow \text{HF}$

In order to be most effective the amount of weak acid and weak base used to prepare the buffer must be considerably greater than the amount strong acid or strong base that may be added later.

Buffer capacity refers to the amount of acid/base that can be neutralized by a buffer solution before its PH is appreciably affected.

PH of buffer solution

To calculate the PH of buffer requires that the concentration of the weak acid (such as HF) and a soluble salt of a weak acid (NaF) be substituted into the ionization constant expression for a weak acid.

For a solution to be buffer ,the ratio of base concentration to the acid concentration shouldnot be less than 0.1 or more than 10.this conclusion follows from handerson Hasselbalchequation ,for a concentration ranging from 1:10 to 10:1implies a pH also ranging

$$PH=Pka+\log 10^{-1} =pka-1 \text{ to } pH=pka+\log 10=pka+1$$

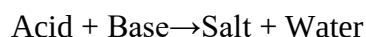
The PH of buffer solution calculate using ***Hhanderson- Hasselbalch*** equation to derive this equation let us consider a buffer made up of weak acid HA and its conjugated base

This phenomena is called maximum buffer capacity (when PH= Pka & POH= Pkb)

7.1.3. Acid–Base Titration and Indicators

Titration is a quantitative analytical technique used to determine the concentration of an unknown solution by reacting it with a solution of known concentration. It is widely used in chemistry, medicine, and industry to analyze substances.

Principle of Titration: Titration involves the slow addition of a titrant (solution of known concentration) to a titrand (solution of unknown concentration) until the reaction reaches its endpoint. The endpoint is indicated by a color change due to an indicator or by an instrumental method. It is determines the concentration of acids or bases using neutralization reactions.



Examples:

Type of Titration	Example Reaction	solution at equivalent point
Strong Acid + Strong Base	$\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$	Neutral (pH=7)
Weak Acid + Strong Base	$\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O}$	Basic (PH>7)
Strong Acid + Weak Base	$\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl}$	Acidic (pH<7)
Weak Acid + Weak Base	$\text{CH}_3\text{COOH} + \text{NH}_3 \rightarrow \text{CH}_3\text{COONH}_4$	

Titration Curve: A titration curve is a graphical representation of pH changes during titration. It helps determine the equivalence point and appropriate indicator.

Acid-base indicators are weak organic acids or weak organic bases that indicate whether a solution is acidic, basic or neutral. Consider a weak organic acid that is denoted by HIn . In order to be effective indicators, HIn , and its conjugate base, In^- , must have different colours. In solution, the acid ionizes as follows: If the indicator is in sufficiently acidic medium, the equilibrium, according to LeChatelier's principle, shifts to the left and the predominant colour of the indicator is that of nonionized form (HIn). This shows us how the ratio of $[\text{In}^-]/[\text{HIn}]$ varies with the concentration of hydronium ion.

- ❖ Many different substances can be used as indicators, depending on the particular reaction to be monitored. For example, red cabbage juice contains a mixture of colored substances that change from deep red at low pH to lightblue at intermediate pH to yellow at high pH.
- ❖ In all cases, though, a good indicator must have the following properties:
 - ✓ The color change must be easily detected.
 - ✓ The color change must be rapid.
 - ✓ The indicator molecule must not react with the substance being titrated.
 - ✓ To minimize errors, the indicator should have a pK_{in} that is within one pH unit of the expected pH at the equivalence point of the titration.

Indicator	pH Range	Color Change
Phenolphthalein	8.2-10.0	Colorless to Pink
Methyl Orange	3.1-4.4	Red to Yellow
Bromothymol Blue	6.0-7.6	Yellow to Blue

- ✓ The choice of an indicator depends on the type of titration and the expected pH at the equivalence point. example:
 - * Strong Acid-Strong Base: Phenolphthalein or Bromothymol Blue
 - * Weak Acid-Strong Base: Phenolphthalein
 - * Strong Acid-Weak Base: Methyl Orange
 - * Weak Acid-Weak Base: No suitable indicator, pH meter recommended

Applications of Titration

- ✓ Determining the concentration of unknown acids/base
- ✓ Quality control in industries (e.g., food, pharmaceuticals, water treatment)
- ✓ Medical applications (e.g., blood acidity analysis)
- ✓ Environmental monitoring (e.g., water hardness determination)

7.1.1. Question with Explanation

271. Which indicator is suitable for titration of HCl with NaCl
A. Phenolphthalein B. Methyl Orange C. Alizarian Red D. Methyl red

Solution

- Titration is a chemical technique used to determine the concentration of a given analyte in a mixture.
- Titration of HCl and NaOH is an example of acid-base titration, which proceeds through neutralisation.
- To determine end point of a titration, indicators are used depending upon what is the strength of acid and base used.
- Both HCl and NaOH being strong acid and strong base respectively, an indicator like **phenolphthalein** can be used which becomes colourless in acidic and neutral conditions and pink coloured under alkaline conditions.

Titration	Titrand	Titrant	Suitable indicators
SA-SB	SA	SB	Phenolphthalien (mostly) Methyl Orange
WA-SB	WA	SB	Phenolphthalien
WB-SA	WB	SA	Methyl Orange

272. Which is the best acid to use to make a buffer solution with a pH = 1.98

- A. Acetic acid (pKa = 4.74) C. chlorous acid (pKa = 1.95)
B. Formic acid (pKa = 3.74). D. Amoonia (pKb=4.74)

Brief:

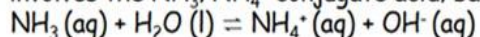
- ✓ To ensure high capacity, we need $[HA] / [A^-]$. It is possible to get this and have a pH = 2.0 only if we use chlorous acid since its pKa is close to the target pH we need.
- ✓ The answer is **chlorous acid (HClO₂)** is best buffer capacity while others are least buffer capacity

273. A buffer is prepared from 60.0 mL of 0.100 M NH_3 and 40.0 mL of 0.100 M NH_4Cl . What is the pH of the buffer solution? $K_b = 1.8 \times 10^{-5}$ for NH_3 .

- A. 4.57 B. 9.43 C. 2.57 D. 4.222

Explanation

The buffer involves the $\text{NH}_3/\text{NH}_4^+$ conjugate acid/base pair.



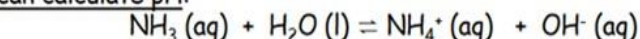
Conc's: Since two solutions are mixed, both have been diluted.

Use this: $V_i M_i = V_f M_f$ $V_f = 60.0 \text{ mL} + 40.0 \text{ mL} = 100.0 \text{ mL}$

$$[\text{NH}_3]_f = V_i M_i / V_f = (60.0 \text{ mL})(0.100 \text{ M}) / 100.0 \text{ mL} = 0.0600 \text{ M}$$

$$[\text{NH}_4^+]_f = V_i M_i / V_f = (40.0 \text{ mL})(0.100 \text{ M}) / 100.0 \text{ mL} = 0.0400 \text{ M}$$

Now, we can calculate pH:



$$[\text{init}] \quad 0.0600 \quad \quad 0.0400 \quad \quad 0$$

$$[\text{change}] \quad -x \quad \quad +x \quad \quad +x$$

$$[\text{equil}] \quad (0.0600-x) \quad \quad (0.0400+x) \quad \quad x$$

$$K_b = 1.8 \times 10^{-5} = \frac{x(0.0400+x)}{(0.0600-x)} \approx \frac{0.0400x}{0.0600} \quad \therefore x = 2.7 \times 10^{-5} \text{ M}$$

$$\therefore \text{pOH} = -\log(2.7 \times 10^{-5}) = 4.57 \quad \therefore \text{pH} = 9.43$$

If we now add strong acid or base, we would again have to calculate new total volume and new $[\text{NH}_3]$ and $[\text{NH}_4^+]$ and [acid or base] to calculate new pH.

274. Which of the following conditions are Not feasible to maximum buffer capacity

- A. $[\text{HA}]$ and $[\text{A}^-]$ are large
 B. $[\text{HA}] = [\text{A}^-]$
 C. The pH is equal to (or very near) the pK_a of the HA used.
 D. The pH is equal to the pOH of buffer solution

Brief:

- Buffer Capacity is how much added acid or base a to a solution can before it can no longer resist changes in the pH.
- The more concentrated the components of the buffer, the greater the buffer capacity
- The more similar the concentrations of the two components, the higher are the buffer capacity.
- The greater the difference, the greater will be the pH change when strong acid or base is added
- Buffer capacity highest when $[\text{HA}] = [\text{A}^-]$
- $\text{pH} = \text{pK}_a$ gives the highest buffer capacity

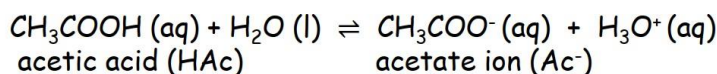
$$\text{pH} = \text{pK}_a + \log \frac{[\text{A}^-]}{[\text{HA}]} \quad \therefore \text{when } [\text{A}^-] = [\text{HA}], \quad \log \frac{[\text{A}^-]}{[\text{HA}]} = \log 1 = 0$$

275. Which of the following is the Not Net Effect for addition of sodium acetate

(CH_3COONa or NaAc) to acetic acid (CH_3COOH or HAc) solution

- A. $[\text{H}_3\text{O}^+]$ decreases C. pH decrease
 B. $[\text{OH}^-]$ decrease D. pOH increase

Brief:



← if we add CH₃COO⁻
SHIFT

Net Effect: [H₃O⁺] decreases ∴ pH increases
Also, [OH⁻] increases ∴ pOH decreases

276. Calculate the pH of the solution that results when 40.0 mL of 0.100 M NH₃ is diluted with 20.0 mL of distilled water.

A. 11 B. 3 C. 4 D. 10

Brief:

➤ Use the dilution equation:

- ❖ $M_1V_1 = M_2V_2$
- ❖ $(0.100 \text{ mol/L})(0.040 \text{ L}) = (x)(0.060 \text{ L})$
- ❖ $x = 0.067 \text{ M}$

➤ Determine pOH of solution using ICE method.

➤ Use the Kb for ammonia to determine pH:

	$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$		
Initial:	0.067	0	0
Change:	-x	x	x
Equilibrium:	0.067-x	x	x

- ❖ $K_b = ([\text{NH}_4^+][\text{OH}^-]) / [\text{NH}_3]$
- ❖ $1.77 \times 10^{-5} = (x)(x) / (0.067 - x)$
- ❖ Remember to ignore the x in 0.067 - x.
- ❖ $x = 0.001 \text{ M}$
- ❖ $\text{pOH} = -\log 0.001 = 3, \text{pH} = 11$

277. How many grams of ammonia are needed to make 1.25 L solution with a pH of 11.68?
- A. 27.6 B. 2.76 C. 276 D. 3.76

Brief:

- Use the pH to get the hydroxide ion concentration:
 - ✓ $\text{pH} = 11.68$
 - ✓ $\text{pOH} = 14.00 - 11.68 = 2.32$
 - ✓ $[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-2.32}$
 - ✓ $[\text{OH}^-] = 4.7863 \times 10^{-3}$
- Use the K_b expression to get the $[\text{NH}_3]$:
 - ✓ $K_b = ([\text{NH}_4^+][\text{OH}^-]) / [\text{NH}_3]$
 - ✓ $1.77 \times 10^{-5} = (4.7863 \times 10^{-3})(4.7863 \times 10^{-3}) / (x)$
 $x = 1.294275 \text{ M}$
- Grams needed for 1.25 L
 - ✓ $MV = \text{Given mass/molar mass}$
 - ✓ $(1.294275 \text{ mol/L})(1.25 \text{ L}) = \text{Given mass}/17.031 \text{ g/mol}$
Given mass = 27.6

278. Which of the following is the pH of a formic acid solution that contains 1.45% formic acid by mass. (Assume a density of 1.01 g/mL for the solution.)
- A. 2.12 B. 3.34 C. 4.21 D.

Brief:

- ✓ Calculate volume of 100.0 g of solution:
- ✓ $100.0 \text{ g divided by } 1.01 \text{ g/mL} = 99.01 \text{ mL}$
- ✓ Calculate molarity of 1.45% solution of HCOOH :
- ✓ $1.45 \text{ g divided by } 46.025 \text{ g/mol} = 0.0315 \text{ mol}$
- ✓ $0.0315 \text{ mol} / 0.09901 \text{ L} = 0.318 \text{ M}$
- ✓ Calculate $[\text{H}^+]$ in 0.318 M HCOOH solution:
- ✓ $1.8 \times 10^{-4} = [(x)(x)] / 0.318$
- ✓ $x = 7.5657 \times 10^{-3} \text{ M}$ (kept a couple guard digits)
- ✓ PH of solution can be calculated as
- ✓ $-\log 7.5657 \times 10^{-3} = 2.12$

279. 4 M KHCO_3 and 0.239 M Na_2CO_3 . If 0.0464 moles of potassium hydroxide are added to 225.0 mL of this buffer, what is the pH of the resulting solution ? (Assume that the volume does not change upon adding potassium hydroxide.)
- A. 10.7 B. 11.2 C. 5.3 D. 3,8

Brief

- moles of bicarbonate and carbonate:
 - ✓ HCO_3^- : (0.384 mol/L) (0.2250 L) = 0.0864 mol
 - ✓ CO_3^{2-} : (0.239 mol/L) (0.2250 L) = 0.053775 mol
- The hydroxide reacts with the acid (the bicarbonate):
 - ✓ HCO_3^- decreases: $0.0864 - 0.0464 = 0.0400$ mol
 - ✓ CO_3^{2-} increases: $0.053775 + 0.0464 = 0.100175$ mol
- pKa of bicarbonate:
 - ✓ Ka of HCO_3^- is the same as the Ka_2 of H_2CO_3
 - ✓ Ka of $\text{HCO}_3^- = 4.7 \times 10^{-11}$
 - ✓ pKa = 10.252
- Use Henderson-Hasselbalch Equation:
 - ✓ $\text{pH} = 10.328 + \log (0.100 / 0.04)$
 - ✓ $\text{pH} = 10.328 + 0.398 = 10.726$

280. How many mL of 4.50 M sodium hydroxide must be added to 250.0 mL of a 0.200 M acetic acid solution to make a buffer with pH = 5.0?

Brief

- Use Henderson-Hasselbalch Equation:
 - ✓ $5.000 = 4.752 + \log [\text{base}] / [\text{acid}]$
 - ✓ $\log [\text{base}] / [\text{acid}] = 0.26$
 - ✓ $[\text{base}] / [\text{acid}] = 1.8$
- Determine molar amount of base required to get pH = 5.000 (for convenience, I'm going to use 1.00 L. I'll go to 250 mL at the end of this step):
 - $1.8x + x = 0.200$

$x = 0.07$ mol (this is acetic acid needed in the solution)

$0.200 - 0.07 = 0.1278$ mol of base required
 $0.1278 / 4 = 0.03195$ mol of acetate required in 250 mL

3) Determine volume of NaOH solution required:

4) $50 \text{ mol/L} = 0.03195 \text{ mol} / x$

$x = 0.0071 \text{ L} = 7.1 \text{ mL}$

4) Check everything:

$[\text{acetic acid}] = 0.01805 \text{ mol} / 0.2571 \text{ L} = 0.070206 \text{ M}$

$[\text{acetate}] = 0.03195 \text{ mol} / 0.2571 \text{ L} = 0.12427 \text{ M}$

$\text{pH} = 4.752 + \log (0.12427 / 0.070206)$

$\text{pH} = 4.752 + \log 1.77$

$\text{pH} = 4.752 + 0.248 = 5.000$

281. If in a solution $10[\text{H}_3\text{O}^+] = [\text{OH}^-]$ then the pH and pOH of the solution are _____ and _____ respectively:
A. 7.5 and 6.5 B. 6.5 and 7.5 C. 2.5 and 11.5 D. 5.5 and 7.5

Brief:

➤ Let: $[\text{H}_3\text{O}^+] = m$ $[\text{OH}^-] = n$

➤ Two equations in two unknowns:

$$10m = n \text{ and } mn = 1.00 \times 10^{-14}$$

➤ Rearrange the second equation:

$$n = (1.00 \times 10^{-14}) / m$$

➤ Substitute into the first equation:

$$10m = (1.00 \times 10^{-14}) / m$$

➤ Rearrange and solve:

$$10m^2 = 1.00 \times 10^{-14}$$

$$m^2 = 1.00 \times 10^{-15}$$

$$m = 3.16 \times 10^{-8} \text{ M}$$

$$m = [\text{H}_3\text{O}^+] = 3.16 \times 10^{-8} \text{ M}$$

➤ Compute the pH:

$$\text{pH} = -\log 3.16 \times 10^{-8} = 7.5$$

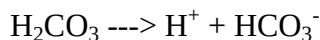
$$\text{pOH} = 14 - \text{pH} =$$

$$= 14 - 7.5 = 6.5$$

282. The solubility of carbonic acid, H_2CO_3 in pure water is 0.0037 mol/L . What is the pH of a 0.0037 M solution of dissolved H_2CO_3 ? K_{a1} for $\text{H}_2\text{CO}_3 = 4.3 \times 10^{-7}$

Brief:

The relevant chemical equation is:



➤ Substituting into the K_a expression, we find

$$4.3 \times 10^{-7} = [(x)(x)] / 0.0037$$

$$x = 3.989 \times 10^{-5} \text{ M}$$

3) Determine the pH:

$$\text{pH} = -\log [\text{H}^+] = -\log 3.989 \times 10^{-5}$$

$$\text{pH} = 4.40 \text{ (to two sig figs)}$$

283. A 0.484 g sample of impure NH_4Br is treated with 25.00 mL of 0.2050 M NaOH and heated to drive off the NH_3 . The unreacted NaOH in the reaction mixture after heating required 9.95 mL of 0.05 M $\text{H}_2\text{C}_2\text{O}_4$ to neutralize. How many grams of NH_4Br were in the original sample?

Brief:

- Determine moles of NaOH added to ammonium bromide solution:

$$M=n/V, n=M \times V(L)$$

$$(0.2050 \text{ mol/L}) (0.02500 \text{ L}) = 5 \times 10^{-3} \text{ mol}$$

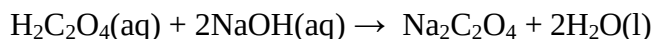
Keep in mind that only some of this NaOH reacted.

- Determine moles of oxalic acid that reacted with excess NaOH :

$$(0.05 \text{ mol/L}) (0.00995 \text{ L}) = 5.2 \times 10^{-4} \text{ mol}$$

- Determine moles of unreacted NaOH :

(use mole mole ratio)



- Demonstrates a 1:2 molar ratio between oxalic acid and NaOH .

$$5 \times 10^{-4} \text{ mol of oxalic acid} \times 2 \text{ mole of NaOH/1mole of oxalic acid}$$

$$= 1 \times 10^{-3} \text{ mol of NaOH reacted}$$

- Determine the amount of NaOH that did not react with the ammonium bromide.

$$\text{Unreacted} = \text{initial mole NaOH} - \text{reacted mole of NaOH}$$

$$5 \times 10^{-3} \text{ mol} - 1 \times 10^{-3} \text{ mol} = 4 \times 10^{-3} \text{ mol}$$

- 5) Determine grams of NH_4Br :

$$\text{Mass of NH}_4\text{Br} = \text{mole of NH}_4\text{Br} \times \text{Molar mass of (NH}_4\text{Br} = \text{NH}_4\text{Br})$$

$$4 \times 10^{-3} \text{ mol} \times 97.943 \text{ g/mol} = 0.400 \text{ g}$$

284. Which of the following combinations cannot produce a buffer solution?

A. HNO_2 and NaNO_2

C. HClO_4 and NaClO_4

B. HCN and NaCN

D. NH_3 and $(\text{NH}_4)_2\text{SO}_4$

Brief:

- A buffer solution consists of a mixture
 - Weak acid and its conjugated base/salt
 - Weak base and its conjugated acid/salt
 - Neither strong acid nor strong base are used in the formation of buffer solution.eg.
 - ✓ HNO_2 (Weak acid/s / HNO_2^- (strong Conjugate base).....buffer solution
 - ✓ HCN (weak acid) / CN^- (conjugated base).... buffer solution

- ✓ NH_3 (Weak base) / NH_4^+ (Conjugated acid) buffer solution
- ✓ HClO_4 (strong acid) and NaClO_4 (weak base).....Not buffer solution

285. What is the pH of a solution composed of 0.20 M NH_3 and 0.15 M NH_4Cl ?

- A. 2.15 B. 4.62 C. 8.26 D. 9.38 E. 8.89

Brief:

What is the pH of a solution containing 0.20 M NH_3 , and 0.15 M NH_4Cl ?

This will be a basic buffer - you are left with the choice of 8.26 or 9.38

- $K_b \text{ NH}_3 = 1.8 \times 10^{-5}$ and $\text{p}K_b = 4.74$
- Calculate pOH
- $\text{pOH} = \text{p}K_b - \log ([\text{NH}_4\text{Cl}] / [\text{NH}_3])$
- $\text{pOH} = 4.74 - \log (0.2/0.15)$
- $\text{pOH} = 4.74 - \log 1.33$
- $\text{pOH} = 4.74 - 0.12$
- $\text{pOH} = 4.62$
- $\text{pH} = 14.00 - 4.62$
- $\text{pH} = 9.38$

286. Calculate the ratio $[\text{CH}_3\text{COOH}]/[\text{NaCH}_3\text{COO}]$ that gives a solution with $\text{pH} = 5.00$?

- A. 0.28 B. 0.36 C. 0.44 D. 0.56 E. 0.63

Brief:

➤ start by using the Henderson - Hasselbalch equation:

- $\text{pH} = \text{p}K_a + \log ([\text{CH}_3\text{COONa}] / [\text{CH}_3\text{COOH}])$
- $\text{p}K_a = -\log (1.8 \times 10^{-5}) = 4.74$
- $5.00 = 4.74 + \log ([\text{CH}_3\text{COONa}] / [\text{CH}_3\text{COOH}])$
- $\log ([\text{CH}_3\text{COONa}] / [\text{CH}_3\text{COOH}]) = 5.00 - 4.74$
- $\log ([\text{CH}_3\text{COONa}] / [\text{CH}_3\text{COOH}]) = 0.26$
- $([\text{CH}_3\text{COONa}] / [\text{CH}_3\text{COOH}]) = 10^{0.26}$
- $([\text{CH}_3\text{COONa}] / [\text{CH}_3\text{COOH}]) = 1.82$
- The mol ratio $\text{CH}_3\text{COONa} : \text{CH}_3\text{COOH}$ must be 1.82 : 1
- The mole ratio of $\text{CH}_3\text{COOH} : \text{CH}_3\text{COONa}$ is $1:1.82=0.56$

287. Consider a solution which is 0.10 M in CH_3COOH and 0.20 M in NaCH_3COO . Which of the following statements is true?

- A. If a small amount of NaOH is added, the pH decreases very slightly.
- B. If NaOH is added, the OH^- ions react with the CH_3COO^- ions.

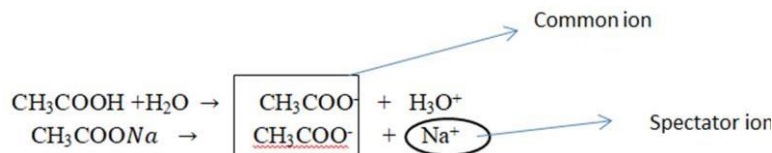
C. If a small amount of HCl is added, the pH decreases very slightly.

D. If HCl is added, the H^+ ions react with CH_3COOH ions.

E. If more CH_3COOH is added, the pH increases.

Brief:

- For solution



- On addition of small amount of CH_3COONa (strong ionized),
- ✓ The $[\text{CH}_3\text{COO}^-]$ increase
 - ✓ According to Lechatlier's principle the equilibrium shift to left
 - ✓ The $[\text{CH}_3\text{COO}^-]$ increased, will consuming the more (H_3O^+) .
 - ✓ Ionization of acetic acid would suppressed/limited.
- On Addition of NaOH:
- ✓ OH^- added is reacting with CH_3COOH Not with CH_3COO^-
 - ✓ $\text{CH}_3\text{COOH} + \text{OH}^- \rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O}$ that suppress the change of pH
 - ✓ The PH increase very slightly (almost suppressed change of pH)
- On Addition of HCl:
- ✓ H_3O^+ added is reacting with CH_3COO^- , Not with CH_3COOH .
 - ✓ $\text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+ \rightarrow \text{CH}_3\text{COOH} + \text{H}_2\text{O}$
 - ✓ The pH decrease very slightly (almost suppressed change of pH)

288. A buffer is prepared by mixing 0.90 mole of ammonia and 0.85 mole of ammonium chloride to form an aqueous solution with a total volume of 0.45 litre. 30.0 ml of 1.00 M NaOH was added to 300 ml of this solution. What is the pH of the solution before and after the addition of NaOH respectively?

A. 4.73 and 4.74 B. 0.93 and 0.92 C. 9.28 and 9.3 D. 9.3 and 9.28

Brief

- The react ion that would occur is as follows- $\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$
- Given-
- The number of moles of ammonia (NH_3) = 0.90 mol
 - The number of moles of ammonium chloride (NH_4Cl) = 0.85 mol
 - The volume of solution = 0.45 L
 - The concentration of NH_3 = Number of moles of NH_3 / volume (L)
= (0.90 mol) / (0.45 L)
= 2 mol / L
 - The concentration of NH_4Cl = (0.85 mol) / (0.45 L)
= 1.89 mol / L

➤ Using Henderson- Hasselbalch equation,

- $pOH = pK_b + \log \frac{[salt]}{[base]}$
- $pOH = pK_b + \log \frac{[salt]}{[base]}$
- $pOH = 4.74 + \log (1.89) / (2)$
- $pOH = 4.74 - 0.02$
- The pH of the solution before the addition of NaOH = 9.28
- The concentration of NaOH = 1.00 M = 1.00 mol / L
- The volume of NaOH solution added = 30.0 mL = 0.03 L
- The number of moles of NaOH = concentration $\times$ Volume

$$= (1.00 \text{ mol / L}) \times (0.03 \text{ L})$$

$$= 0.03 \text{ mol}$$

- The given volume of the solution in which NaOH solution is added = 300 mL = 0.3 L
- The total volume of the solution becomes (0.3 L + 0.03 L) = 0.33 L

➤ After the addition of NaOH (Base) solution, the concentration becomes as-

On addition of 0.03 mole of OH^- (from NaOH), 0.03mole of NH_4^+ consumed and provides 0.03 mole of NH_3 (New).

- The concentration of NH_3 = $\frac{\text{moles of } NH_3 + \text{moles of NaOH}}{\text{total volume of solution}}$
 $= (0.90 \text{ mol} + 0.03 \text{ mol}) / (0.33 \text{ L})$
 $= 2.8 \text{ mol / L}$

- The concentration of NH_4Cl = $\frac{\text{moles of } NH_4Cl - \text{moles of NaOH}}{\text{total volume of solution}}$
 $= (0.85 \text{ mol} - 0.03 \text{ mol}) / (0.33 \text{ L})$
 $= 2.5 \text{ mol / L}$

➤ Using Henderson- Hasselbalch equation,

- $pOH = pK_b + \log ([salt] / [base])$
- $pOH = 4.74 + \log (2.5) / (2.8)$
- $pOH = 4.74 - 0.05$
- $pOH = 4.7$
- $pH + pOH = 14$
- $pH = 14 - 4.7$
- $pH = 9.3$
- The pH of the solution after the addition of NaOH solution = 9.3
- Change of pH = pH after addition of NaOH – pH before addition of NaOH
- $\Delta pH = 9.3 - 9.28 = 0.02$
- *Therefore the change in pH is resist on addition of NaOH solution.*

289. How many grams of NaF would have to be added to 2.00 L of 0.100 M HF to yield a solution with a pH = 3.5?

- A. 1.93 g B. 193 g C. 84 g D. 19.3 g

Brief

- Using Henderson Hasselbalch equation:
 - ✓ $\text{pH} = \text{pK}_a + \log [\text{salt}] / [\text{acid}]$
 - ✓ $\text{pH} = 3.50$
 - ✓ $\text{pK}_a = -\log K_a$ and K_a for HF = 7.2×10^{-4} ; $\text{pK}_a = 3.14$
 - ✓ $[\text{HF}] = 0.100 \text{ M}$
 - ✓ $[\text{NaF}] = ?$
- So, we will first solve for $[\text{NaF}]$ and then determine grams needed:
 - ✓ $3.50 = 3.14 + \log (\text{NaF}/0.100)$
 - ✓ $0.36 = \log (\text{NaF}/0.100)$
 - ✓ $[\text{NaF}] / 0.100 = 2.3$
 - ✓ $[\text{NaF}] = 0.23 \text{ M}$
 - ✓ Molar mass NaF = 41.99 g / mol
 - ✓ $\text{Molarity} = \frac{\text{Given mass}}{\text{Molar mass} \times \text{volume (L) of solution}}$
 - ✓ $\text{Given mass} = \text{Molarity} \times \text{molar mass} \times \text{volume (L)}$
 - ✓ $0.23 \text{ mol / L} \times 2 \text{ L} \times 41.99 \text{ g / mol} = 19.3 \text{ g NaF needed.}$

290. Calculate the pH that results when the following solutions are mixed.

- (1) 35 mL of 0.20 M formic acid
 (2) 55 mL of 0.10 M sodium formate
 (3) 110 mL of water

- A. 3.6 B. 3.1 C. 4.58 D. 3.39 E. 4.20

Brief:

- Given
 - ✓ Mole of formic acid = $35 \text{ mL} (0.035 \text{ L}) \times 0.2 \text{ mol/L} = 0.007 \text{ mol}$
 - ✓ Mole of sodium formate = $55 \text{ mL} (0.055 \text{ L}) \times 0.1 \text{ mol/L} = 0.0055 \text{ mol}$
 - ✓ Total volume of solution = $35 \text{ mL} + 55 \text{ mL} + 110 \text{ mL} = 200 \text{ mL} = 0.2 \text{ L}$
 - ✓ The number of mole of HCOOH originally = 0.007 mol
 - ✓ The number of mole of HCOONa originally = 0.0055 mol
- Molarity of HCOOH in solution
 - = mole of HCOOH / total volume of solution
 - = $0.007 \text{ mol} / 0.2 \text{ L}$
 - = 0.035 M
- Molarity of HCOO^-
 - = mole of HCOO^- / total volume of solution
 - = $0.0055 \text{ mol} / 0.2 \text{ L}$
 - = 0.0275 M

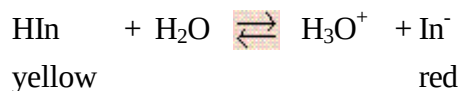
- ✓ Hydrolysis of formic acid



	HCOOH	HCOO ⁻	H ₃ O ⁺
Initial (M)	0.035	0.03	0
Change	-X	X	X
equilibrium	0.035-X	0.03+X	X

- ✓ $K_a = [\text{HCOO}^-][\text{H}_3\text{O}^+]/[\text{HCOOH}]$
- ✓ $1.8 \times 10^{-4} = (X)(0.03+X)/(0.035-X)$
- ✓ Solving X, and $X = 2.1 \times 10^{-4}$ (using approximation method. check its validity)
- ✓ $X = \text{H}_3\text{O}^+ = 2.1 \times 10^{-4}$
- ✓ $\text{pH} = -\text{Log} (2.1 \times 10^{-4})$
- ✓ $\text{pH} = 3.6$

291. Consider an indicator that ionized as shown below for which its $K_a = 1.0 \times 10^{-4}$



Which of the responses contains all the true statements and no others?

- (1) The predominant color in its acid range is yellow.
 - (2) In the middle of the pH range of its color change a solution containing the indicator will probably be orange.
 - (3) At $\text{pH} = 7.00$, a solution containing this indicator (and no other colored species) will be red. (Hint: Write the equilibrium constant expression for the indicator.)
 - (4) At $\text{pH} = 7.00$, most of the indicator is in the un-ionized form.
 - (5) The pH at which the indicator changes color is $\text{pH} = 4$.
- A. 1, 3, 5 B. 2, 4 C. 3, 4, 5 D. 1, 2, 3, 5

Brief:

$$K_{\text{In}} = \frac{[\text{In}^-][\text{H}_3\text{O}^+]}{[\text{HIn}]}$$

- ✓ The ratio of $\text{HIn} : \text{In}^- = [\text{H}_3\text{O}^+]/K_{\text{In}}$
- ✓ The color of solution attains the color of acidic form of indicator (yellow) at $\text{pH} < 4$.
- ✓ The color of solution attains the color of basic form of indicator (red) at $\text{pH} > 4$.
- ✓ The color of solution attains the mixture of yellow and red (orange) at $\text{pH} = 4$.
- ✓ The pH at which the indicator changes color is start at $\text{pH} = 4$.

292. Calculate the pH of the solution resulting from the addition of 20.0 mL of 0.100 M NaOH to 30.0 mL of 0.100 M HNO₃. A. 1.35 B. 1.70 C. 1.95 D. 2.52 E. 2.80

Brief:

- This is strong acid strong base titration (HNO₃) titrated by NaOH)
- Initial moles of NaOH = (0.100 mol/L) × (20.0/1000 L) = 0.002 mol
- Initial moles of HNO₃ = (0.100 mol/L) × (30.0/1000 L) = 0.003 mol
- Balanced equation for the reaction: NaOH + HNO₃ → NaNO₃ + H₂O
- Accordingly, mole ratio in reaction NaOH : HNO₃ = 1 : 1
- But initial mole ratio NaOH : HNO₃ = 0.002 : 0.003 = 2 : 3
- Hence, NaOH is the limiting reactant, and HNO₃ is in excess.
 - Moles of NaOH reacted = 0.002 mol
 - Moles of HCl reacted = 0.002 mol
 - 0.002 mol of HCl consuming (neutralizing) 0.002 mol of NaOH
 - Moles of excess HCl remained in the solution
= (0.003 - 0.002) mol = 0.001 mol

Assume that volumes of solution are additive.

Volume of the resulting solution = (20.0 + 30.0) mL = 50.0 mL = 0.05 L

Molarity of HCl in the resulting solution = (0.001 mol) / (0.05 L) = 0.02 M

Since HCl is a strong acid, [H₃O⁺] in the resulting solution = 0.02 M

Hence, **pH of the resulting solution** = -log[H₃O⁺] = -log(0.02) = **1.700**

293. When 100.0cm³ of 0.1 M NaOH is mixed with 100.0cm³ of 0.15 M. What will the concentration of the resulting solution?

A. 0.251M B. 0.187M C. 0.125M D. 0.753M

Brief:

- The key here is to know the total number of moles of dissolved solute and the total volume of the solution in liters:
 - ✓ 100.0 cm³ = 0.1000 L
 - ✓ Mole of NaOH in solution 1= (0.100 M)(0.1000 L) = 0.0100 mol
 - ✓ Mole of NaOH in solution 2= (0.150 M)(0.1000 L) = 0.0150 mol
 - ✓ Total mol = 0.0100 mol + 0.0150 mol = 0.0250 mol
 - ✓ Total volume = 0.1000 L + 0.1000 L = 0.2000 L
 - ✓ Concentration = 0.0250 mol/0.2000 L = 0.125 M OR
 - ✓ The resultant Concentration:
 - M₁V₁+M₂V₂ = M₃V₃
 - M₃ = (M₁V₁+M₂V₂)/V₃
 - M₃= (0.1mol/L × 0.1L) + (0.15mol/L × 0.1L)/(0.1L+0.1L)

294. Consider the titrations of the pairs of aqueous acids and bases listed below. For which pair is the pH at the equivalence point stated incorrectly?

Acid-Base Pair	pH at Equivalence Point
A. $\text{NH}_3 + \text{HCl}$	less than 7
B. $\text{HNO}_3 + \text{NaOH}$	equal to 7
C. $\text{HClO}_4 + \text{NaOH}$	equal to 7
D. $\text{HClO} + \text{NaOH}$	Less than 7
E. $\text{CH}_3\text{COOH} + \text{KOH}$	greater than 7

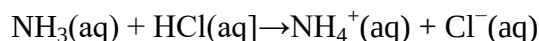
Brief:

Titration	pH ranges at equivalent point	Nature of solution at equivalent point	Suitable indicators	examples
Strong acid – strong base	= 7	Neutral	Phenolphthalein (mostly) Methyl Orange	$\text{HClO}_4 + \text{NaOH}$
Weak acid – strong base	>7	Basic	Phenolphthalein	$\text{CH}_3\text{COOH} + \text{KOH}$
Weak base – strong acid	<7	Acidic	Methyl Orange	$\text{NH}_3 + \text{HCl}$

295. What is the pH at the equivalence point when 35.0 mL of 0.20 M ammonia is titrated by 0.12M hydrochloric acid? K_b for ammonia is 1.8×10^{-5} .

Brief:

- Even without doing any calculations, possible to say that the pH of the solution at equivalence point will be smaller than 7.
- In the case of a weak base being titrated with a strong acid, the neutralization reaction will produce the *conjugate acid* of the weak base, which will then react with water to produce hydronium ions, H_3O^+ , and reform some of the weak base.
- The formation of the hydronium ions will cause the pH of the solution to be smaller than 7 at equivalence point, i.e. the resulting solution will be *acidic*.
- So, ammonia reacts with hydrochloric acid according to the balanced chemical equation



Notice that you have 1:1 mole ratios between the species that take part in the reaction.

- consume one mole of hydrochloric acid for every one mole of ammonia
- produce one mole of ammonium ions for every one mole of ammonia consumed

Use the ammonia solution's molarity and volume to determine how many moles of ammonia you have in the initial solution

$$M=n/V \Rightarrow n=M \cdot V$$

$$n_{\text{NH}_3} = (0.20 \text{ M})(35 \times 10^{-3} \text{ L}) = 0.0070 \text{ moles NH}_3$$

- In order to reach the equivalence point, must add just enough acid to neutralize the base.
- Accordingly, to reach equivalent point the balanced chemical equation, the equivalence point requires equal numbers of moles of ammonia and hydrochloric acid.

This means that you must add, $M=n/V \Rightarrow V=n/M$

- $V_{\text{HCl}} = \frac{0.0070 \text{ moles}}{0.12 \text{ mol/l}}$
- $V_{\text{HCl}} = 0.05833 \text{ L} = 58.3 \text{ mL}$ of hydrochloric acid added to the ammonia solution.

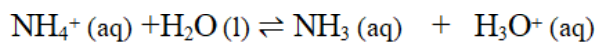
- The **total volume** of the solution will thus be

- $V_{\text{total}} = V_{\text{NH}_3} + V_{\text{HCl}}$
- $V_{\text{total}} = 35.0 \text{ mL} + 58.3 \text{ mL} = 93.3 \text{ mL}$

So, if 0.0070 moles of ammonia and 0.0070 moles of hydrochloric acid were consumed, it follows that the reaction produced 0.0070 moles of *ammonium ions*.

- The **concentration** of the ammonium ions will be

$$[\text{NH}_4^+] = \frac{0.0070 \text{ moles}}{0.0933 \text{ L}} = 0.07503 \text{ M}$$



I:	0.075	0	0
C:	-x	+x	+x
eq:	0.075-x	x	x

- Now, notice that the problem provides you with the *base dissociation constant*, K_b . In order to find the *acid dissociation constant*, K_a , for the ammonium ion, use the equation

$$K_a \cdot K_b = K_w,$$

Where K_w - the self-ionization constant of water, equal to 10^{-14} at room temperature

$$K_a = K_w / K_b = 10^{-14} / 1.8 \cdot 10^{-5} = 5.6 \times 10^{-10}$$

By definition, the acid dissociation constant is equal to

$$K_a = \frac{[\text{NH}_3] \cdot [\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = \frac{x \cdot x}{0.07503 - x}$$

- Since x has such a small value, you can say that, $0.07503 - x \approx 0.07503$ \
- And simplifying of x , $x = 6.5 \cdot 10^{-6}$

Since x represents the equilibrium concentration of the hydronium ions

- ✓ $[\text{H}_3\text{O}^+] = 6.5 \times 10^{-6} \text{ M}$
- ✓ $\text{pH} = -\log([\text{H}_3\text{O}^+])$
- ✓ $\text{pH} = -\log(6.5 \times 10^{-6}) = 5.19$

As predicted, the pH of the resulting solution will be *smaller* than 7 at equivalence point.

296. A 35.0mL sample of an unknown HClO_4 solution requires 50.3mL of 0.109M NaOH for complete neutralization. What was the concentration of the unknown HClO_4 solution? The neutralization reaction is: $\text{HClO}_4(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{NaClO}_4(\text{aq})$

- A. 0.166 M B. 1.66 M C. 2.72 M D. 3.32M

Brief:

- Start with the balanced chemical equation for this neutralization reaction
 $\text{HClO}_4(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{NaClO}_4(\text{aq})$
- Notice the 1:1 mole ratio that exists between the reactants. This tells you that a *complete* neutralization implies equal number of moles of perchloric acid and of sodium hydroxide.
- Use the sodium hydroxide solution's molarity and volume to determine how many moles were needed to completely neutralize the acid
 $M = n/V \Rightarrow n = M \cdot V$, $n(\text{NaOH}) = 0.109 \text{ M} / 50.3 \times 10^{-3} \text{ L} = 0.00548 \text{ moles NaOH}$
- Automatically, this is equal to the number of moles of perchloric acid present in solution. This means that the molarity of the initial perchloric acid solution was

$$M = n/V$$

$$M(\text{HClO}_4) = 0.00548 \text{ moles} / 35.0 \times 10^{-3} \text{ L}$$

297. What is the pH of a solution made by mixing 100.0 mL of 0.10 M HNO₃, 50.0 mL of 0.20 M HCl and 100.0 mL of water? Assume that the volumes are additive.

- A. 1.10 B. 2.67 C. 5.34 D. 11.321

Brief

- First thing first, calculate the total volume of the resulting solution

$$V_{\text{total}} = 100.0 \text{ mL} + 50.0 \text{ mL} + 100.0 \text{ mL}$$

$$V_{\text{total}} = 250.0 \text{ mL}$$

- Now, you are dealing with two strong acids that ionize completely in aqueous solution. Both nitric acid and hydrochloric acid produce hydronium cations in 1:1 mole ratios, so you know that

$$[\text{H}_3\text{O}^+]_{\text{coming from HNO}_3} = [\text{HNO}_3]$$

$$[\text{H}_3\text{O}^+]_{\text{coming from HCl}} = [\text{HCl}]$$

- As you know, molarity is defined as the number of moles of solute present in 1000 mL of solution. For the nitric acid solution, you have

$$100.0 \text{ mL} = \frac{10^3 \text{ mL}}{10} \Rightarrow n_{\text{H}_3\text{O}^+} = \frac{0.10 \text{ moles}}{10} = 0.010 \text{ moles H}_3\text{O}^+$$

For hydrochloric acid, you have

$$50.0 \text{ mL} = \frac{10^3 \text{ mL}}{20} \Rightarrow n_{\text{H}_3\text{O}^+} = \frac{0.20 \text{ moles}}{20} = 0.010 \text{ moles H}_3\text{O}^+$$

- The **total number of moles** of hydronium cations delivered by the two acids in the resulting solution will be

$$n(\text{H}_3\text{O}^+) = 0.010 \text{ moles} + 0.010 \text{ moles}$$

$$n(\text{H}_3\text{O}^+) = 0.020 \text{ moles}$$

- The **concentration** of the hydronium cations in the resulting solution will be

298. Electrolysis of dilute aqueous NaCl solution was carried out by passing 10mA current. This time required to liberate 0.01moles of H₂ gas at the cathode is:

A. $9.65 \times 10^4 \text{ sec}$ C. $19.3 \times 10^4 \text{ sec}$

B. $28.95 \times 10^4 \text{ sec}$ D. $38.6 \times 10^4 \text{ sec}$

Brief:

- We have to calculate the time taken for the liberation of hydrogen gas at the cathode using the amount of hydrogen released (in moles), current passed and Faraday's constant. The amount of hydrogen released in moles is calculated using the electrons transferred during the process multiplied by the moles of hydrogen released. Multiply the amount

of hydrogen released, current passed and Faraday's constant to get time taken for the liberation of hydrogen.

- Current passed is 10mA.

Moles of hydrogen released are 0.01moles.

In the electrolysis of aqueous solution of sodium chloride, as the reduction potential of sodium ion is less than the hydrogen ion, the reaction that takes place at the cathode could be given as,

$$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$$

- 0.01moles of hydrogen would require $2 \times 0.01 = 0.02$ moles of electrons (or) 0.02 faradays.
- We know the formula to calculate the amount of Faradays passed,
 - ✓ $Q = It$
 - ✓ From the equation: 2mole H^+ would require = 96500C charge
 - ✓ 0.01 Mole of H^+ required = ?
 - ✓ Amount of charge passing on discharging of 0.01 mole of H^+ is $0.02 \times 96500\text{C}$
= 1930
 - ✓ Time require is $= Q/I = 1930\text{C}/10^{-2}\text{A}$
 - ✓ $t = 19.3 \times 10^4 \text{sec} = 1.93 \times 10^5 \text{sec}$

299. What is the approximate pH of a solution prepared by mixing equal volumes of 0.05 M methylamine and 0.20 M hydrochloric acid?

A. 2.57 B. 1.12 C. 1.63 D. 10.5 E. 9.8

Brief:

1.125 is the approximate pH of a solution prepared by mixing equal volumes of 0.05 M methylamine and 0.20 M hydrochloric acid.

- ❖ let $V = \text{Volume}$ of both methyl amine = volume of HCl
- ❖ moles of methyl amine = $V \times 0.05 \text{mol/L}$
- ❖ moles of hydrochloric acid = $V \times 0.2 \text{mol/L}$
- ❖ The reaction between methyl amine and HCl is given by
- ❖ $\text{CH}_3\text{NH}_2 + \text{HCl} \rightarrow \text{CH}_3\text{NH}_3^+ + \text{Cl}^-$
- ❖ methyl amine is limiting **reactant** and hydrochloric acid is excess by $= V \times (0.2 - 0.05) = 0.15V$
- ❖ Volume after mixing = $V + V = 2V$
- ❖ Concentration of HCl = $0.15V/2V = 0.075 \text{M}$
- ❖ $\text{pH} = -\log(0.075) = 1.125$

In **chemistry**, pH is a scale used to quantify the acidity or basicity of an aqueous solution. It was traditionally denoted as "potential of hydrogen" (or "power of hydrogen"). Acidic solutions (solutions with greater H^+ ion concentrations) have lower pH values than basic or **alkaline** solutions.

300. Which of the following salts give acidic aqueous solutions?

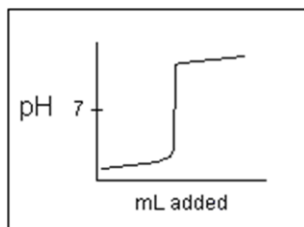
- (1) KNO_3 (2) KCH_3COO (3) NH_4NO_3 (4) RbI
 (5) $(NH_4)_2SO_4$ (6) $BaCl_2$ (7) $NaCN$ (8) KNO_2
 A. 2, 7, 8 B. 3, 5 C. 2, 4, 6 D. 1, 4, 7, 8 E. 1, 4, 6

Brief:

A salt that is derived from the reaction of a strong acid with a strong base forms a solution that has a pH of 7. An example is sodium chloride, formed from the neutralization of HCl by NaOH. A solution of NaCl in water has no acidic or basic properties, since neither ion (Na^+ or Cl^-) is capable of hydrolyzing with H_2O . Other salts that form neutral solutions include potassium nitrate (KNO_3) and lithium bromide ($LiBr$). There are different types of salts, namely acidic salt, basic salt, neutral salt.

Acid	Base	Salt	Hydrolysis	Examples
Strong	Strong	Neutral	No ionic hydrolysis	$NaCl$, KNO_3
Weak	Strong	Basic	anionic hydrolysis	KCN , KNO_2 , CH_3COOK
Strong	weak	acidic	cationic hydrolysis	NH_4NO_3 , $(NH_4)_2 SO_4$
Weak	Weak	Acidic: $K_a > K_b$	Both cationic and anionic hydrolysis	
		Basic: $K_b > K_a$		NH_4CN
		Neutral: $K_a = K_b$		CH_3COONH_4

301. The following titration curve is the kind of curve expected for the titration of a ____ acid with a ____ base.

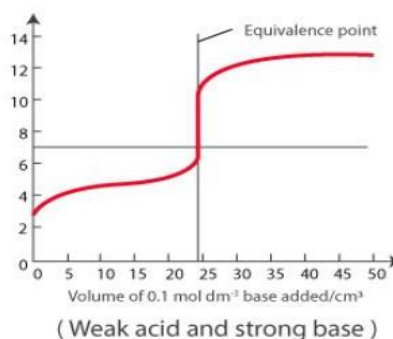
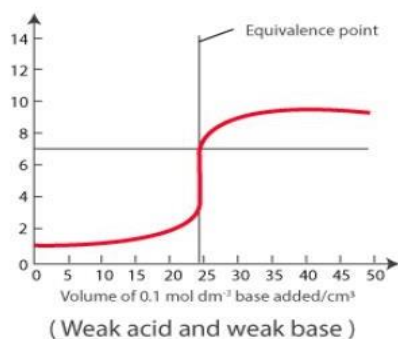
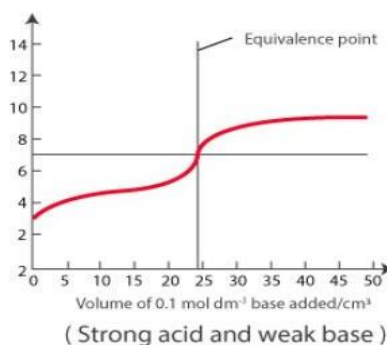
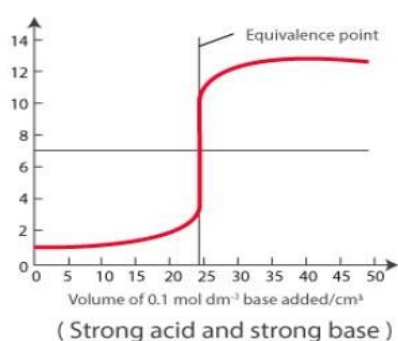


- A. strong, strong
 B. weak, strong
 C. strong, weak
 D. weak, weak

Brief:-

- ❖ Titration – A process where a solution of known strength is added to a certain volume of a treated sample containing an indicator.
- ❖ Titrant – A solution of known strength of concentration used in the titration.
- ❖ Titrand – The titrand is any solution to which the titrant is added and which contains the ion or species being determined.

- ❖ Titration curve – A plot of pH Vs millilitres of titrant showing the manner in which pH changes Vs millilitres of titrant during an acid-base titration.
- ❖ Equivalence point – The point at which just an adequate reagent is added to react completely with a substance.
- ❖ Buffer solution – A solution that resists changes in pH even when a strong acid or base is added or when it is diluted with water.
- ❖ A titration curve is a graphical representation of the pH of a solution during a titration. In a strong acid-strong base titration, the equivalence point is reached when the moles of acid and base are equal and the pH is 7.



302. Consider the titration of 30.0 mL of 0.20 M nitrous acid by adding 0.0500 M aqueous ammonia to it. The pH at the equivalence point is _____. (Note: This is the titration of a weak acid with a weak base.)

- A. greater than 7 B. Equal to 7
C. less than 7 D. cannot be determined without more data, like K_a and K_b .

Brief:-

- ✓ The pH at equivalent of strong Acid and strong base is 7
- ✓ The pH at equivalent of weak Acid and strong base is greater than 7
- ✓ The pH at equivalent of weak base and strong acid is less than 7

303. Arrhenius defined an acid as:

- A. a species that can donate a proton. B. a species that can accept a proton.
 C. a source of OH^- ions in water. D. a source of H^+ ions in water.
304. In the Brønsted-Lowry system, a base is defined as:
 A. a proton donor. B. a hydroxide donor.
 C. an electron-pair acceptor. D. a proton acceptor.

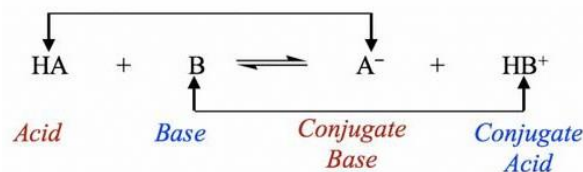
Brief:

Theory	Acids	Bases
Arrhenius	Produce H^+ ions in water	Produce OH^- ions in water
Brønsted-Lowry	Proton donor	Proton acceptors
Lewis	Electron acceptor	Electron donor

305. In the equation: $\text{HF} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{F}^-$
 A. H_2O is a base and HF is its conjugate acid.
 B. H_2O is an acid and HF is the conjugate base.
 C. HF is an acid and F^- is its conjugate base.
 D. HF is a base and H_3O^+ is its conjugate acid.

Brief:

- ❖ The most commonly applied definition of acids and bases is the Brønsted-Lowry definition:
 - ✓ Brønsted-Lowry Acid: a substance that can donate a proton (H^+)
 - ✓ Brønsted-Lowry Base: a substance that can accept a proton (H^+)
- ❖ Therefore, according to the Brønsted-Lowry definition, an acid-base reaction is a proton transfer process in which the acid gives away a proton and the base accepts a proton as shown in the general equation:



The species that forms when an acid loses its proton is called the **conjugate base** of that acid; similarly, the species that forms when a base accepts a proton is called the **conjugate acid** of that base. In the general equation above, HA is the conjugate acid of A^- , and A^- is the conjugate base of HA. HA and A^- can also be called a **conjugate acid-base pair**; another pair is HB^+ and B.



For weak acids (HA is used as a general formula), the proton is only partially donated and the reaction stays at equilibrium. The equilibrium arrow $\rightleftharpoons$ will be needed in the reaction equation to indicate the equilibrium status:



The equilibrium constant for the above reaction is called the acid dissociation constant, K_a . It is a constant to measure the relative strength of an acid. The expression for K_a is:

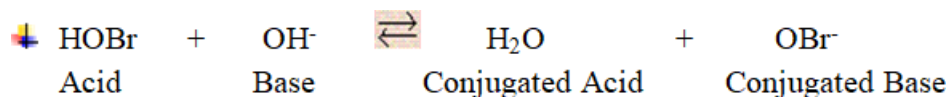
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

The larger the K_a value, the stronger the ability of the acid to donate protons, and the stronger the acid is. (Technically, when the K_a value is larger than 10, the acid can be regarded as a strong acid.)

306. For the system shown here: $\text{HOBr} + \text{OH}^- \rightleftharpoons \text{H}_2\text{O} + \text{OBr}^-$ Bronsted would classify the base species as:

- A. OH^- and HOBr B. H_2O and OH^-
C. OBr^- and OH^- D. OBr^- and HOBr

❖ Given the reaction:-



307. Which is the strongest acid?

- A. HClO_4 B. HClO_3 C. HClO_2 D. HClO

- ✓ Acid strength is the measure of the ability of the acid to lose its H^+ ion.
- ✓ The dissociation of a strong acid in solution is finely complete, omitting in its most concentrated solutions.
- ✓ Examples of strong acids are hydrochloric acid, perchloric acid, nitric acid, and sulfuric acid.

➤ Classification of Acids Based on the Number of Elements they Contain

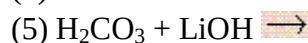
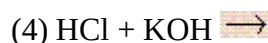
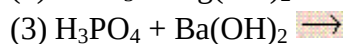
- i. Binary acids: are those acids composed of only two different elements
 - Examples: HCl, HBr, HF, HI, and H_2S .
- ii. Ternary acids: Are acids composed of three different elements.
 - The three elements usually hydrogen, oxygen and a non-metal.

Examples : H_2SO_4 , H_2CO_3 , HClO_4 , and H_3PO_4 . Perchloric acid (HClO_4), chloric acid (HClO_3), chlorous acid (HClO_2), Hypochlorous acid (HClO)

Based on their dissociation

- Strong acids - acid completely dissociates in an aqueous solution.
 - ✓ produce large amount of H^+ ions in the aqueous solution
 - ✓ Their dissociation is indicated by a single arrow ($\rightarrow$) as shown below:
 - ✓ e.g. Hydrochloric acid (HCl), nitric acid (HNO_3), and sulfuric acid (H_2SO_4)
Perchloric acid (HClO_4),
 - $\text{HClO}_4 > \text{HClO}_3$ acid strength
 - $\text{H}_3\text{PO}_2 > \text{H}_3\text{PO}_3 > \text{H}_3\text{PO}_4$ acid strength
- ❖ Weak acids - dissociates only partially
- ❖ Release few hydrogen ions in aqueous solution.
 - Their aqueous solution contains H^+ ion, anions and dissolved molecules of the acid.
 - their dissociation indicated by the double arrow ($\rightleftharpoons$) shown below
 - e.g. Acetic acid (CH_3COOH), oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$), carbonic acid (H_2CO_3).
 - Organic acids, which contain the acidic carboxyl group – COOH , are generally weak acids.
- ❖ Dependence of acid strength
 - ✓ $\text{HF} > \text{H}_2\text{O} > \text{NH}_3 > \text{CH}_4$ – As electronegativity of $\text{F} > \text{O} > \text{N} > \text{C}$.
 - ✓ $\text{HNO}_3 > \text{H}_2\text{CO}_3 > \text{H}_3\text{BO}_3$ - As electronegativity of $\text{N} > \text{C} > \text{B}$
 - ✓ $\text{HClO} > \text{HBrO} > \text{HIO}$ $\text{HClO}_3 > \text{HBrO}_3 > \text{HIO}_3$ As electronegativity of $\text{Cl} > \text{Br} > \text{I}$.
 - ✓ $\text{HF} < \text{HCl} < \text{HBr} < \text{HI}$ - Size of central atom or ease of removal of H^+
 - ✓ $\text{H}_2\text{O} < \text{H}_2\text{S} < \text{H}_2\text{Se}$ - Size of central atom or ease of removal of H^+
 - ✓ $\text{NH}_3 < \text{PH}_3 < \text{AsH}_3 < \text{SbH}_3$ - Size of central atom or ease of removal of H^+
 - ✓ $\text{H}_3\text{PO}_2 > \text{H}_3\text{PO}_3 > \text{H}_3\text{PO}_4$

308. Consider the neutralization reactions between the following acid-base pairs in dilute aqueous solutions:

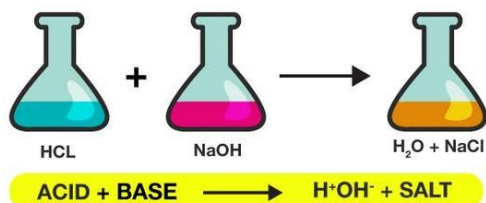


309. For which of the reactions is provides the net ionic equation: $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$?

Brief: A. 1, 3 B. 1, 4, 5 C. 1, 2, 3, 4, 5 E. only 4

✚ A neutralization reaction is when an acid and a base react to form water and a salt and involves the combination of H^+ ions and OH^- ions to generate water.

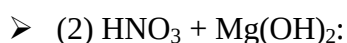
✚ The neutralization of a strong acid and strong base has a pH equal to 7



- When an acid and a base react, they undergo a neutralization reaction. In these reactions, the acid donates a proton (H^+) to the base, forming water and a salt. Let's analyze each of the given acid-base pairs:



In this reaction, acetic acid (CH_3COOH) is an acid, while sodium hydroxide (NaOH) is a base. The hydrogen ion (H^+) from the acid combines with the hydroxide ion (OH^-) from the base to form water (H_2O). The remaining sodium ion (Na^+) from the base combines with the acetate ion (CH_3COO^-) from the acid to form sodium acetate (NaCH_3COO).



- In this reaction, nitric acid (HNO_3) is an acid, while magnesium hydroxide ($\text{Mg}(\text{OH})_2$) is a base. The hydrogen ion (H^+) from the acid combines with the hydroxide ion (OH^-) from the base to form water (H_2O). The remaining magnesium ion (Mg^{2+}) from the base combines with the nitrate ion (NO_3^-) from the acid to form magnesium nitrate ($\text{Mg}(\text{NO}_3)_2$).

- $\text{H}_3\text{PO}_4 + \text{Ba}(\text{OH})_2$:

In this reaction, phosphoric acid (H_3PO_4) is an acid, while barium hydroxide ($\text{Ba}(\text{OH})_2$) is a base. The hydrogen ion (H^+) from the acid combines with the hydroxide ion (OH^-) from the base to form water (H_2O). The remaining barium ion (Ba^{2+}) from the base combines with the phosphate ion (PO_4^{3-}) from the acid to form barium phosphate ($\text{Ba}_3(\text{PO}_4)_2$).

- $\text{HCl} + \text{KOH}$:

In this reaction, hydrochloric acid (HCl) is an acid, while potassium hydroxide (KOH) is a base. The hydrogen ion (H^+) from the acid combines with the hydroxide ion (OH^-) from the base to form water (H_2O). The remaining potassium ion (K^+) from the base combines with the chloride ion (Cl^-) from the acid to form potassium chloride (KCl).

- $\text{H}_2\text{CO}_3 + \text{LiOH}$:

In this reaction, carbonic acid (H_2CO_3) is an acid, while lithium hydroxide (LiOH) is a base. The hydrogen ion (H^+) from the acid combines with the hydroxide ion (OH^-) from the base to form water (H_2O). The remaining lithium ion (Li^+) from the base combines with the bicarbonate ion (HCO_3^-) from the acid to form lithium bicarbonate (LiHCO_3).

In all of these reactions, water and a salt are formed. The salt consists of the positive ion from the base and the negative ion from the acid. It's important to note that the formation of water in these reactions helps to neutralize the acidic and basic properties of the reactants, resulting in a neutral solution.

310. Which one of the following represents the net ionic equation for the reaction of nitric acid with aluminum hydroxide?

- A. $3\text{H}^+ + \text{Al}(\text{OH})_3 \rightarrow \text{Al}^{3+} + 3\text{H}_2\text{O}$
B. $3\text{HNO}_3 + \text{Al}(\text{OH})_3 \rightarrow \text{Al}(\text{NO}_3)_3 + 3\text{H}_2\text{O}$
C. $\text{HNO}_3 + \text{OH}^- \rightarrow \text{NO}_3^- + \text{H}_2\text{O}$
D. $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
E. $3\text{NO}_3^- + \text{Al}^{3+} \rightarrow \text{Al}(\text{NO}_3)_3$

Brief:

The net ionic equation is the chemical equation that shows only those elements, compounds, and ions that are directly involved in the chemical reaction.

The net ionic equation for the reaction of aluminum hydroxide with hydrochloric acid is $\text{Al}(\text{OH})_3(\text{s}) + 3\text{H}^+(\text{aq}) \rightarrow \text{Al}^{3+}(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$, highlighting the transformation of solid aluminum hydroxide into dissolved aluminum ions and water.

311. Which one of the following is an amphoteric metal hydroxide?
A. KOH B. $\text{Ba}(\text{OH})_2$ C. $\text{Pb}(\text{OH})_2$ D. LiOH

Brief:

Amphoteric hydroxides are those compounds that can react with both acids and alkalis to produce salt and water. The hydroxides of zinc, aluminium, lead, etc. are amphoteric in nature. Examples of amphoteric hydroxides are $\text{Zn}(\text{OH})_2$, $\text{Pb}(\text{OH})_2$, and $\text{Al}(\text{OH})_3$

General notation:

1. Amphoteric hydroxide + acid $\rightarrow$ salt + water
$$\text{Al}(\text{OH})_3 + 3 \text{HCl} \rightarrow \text{AlCl}_3 + 3\text{H}_2\text{O}$$
2. Amphoteric hydroxide + base $\rightarrow$ hydroxo-complex (salt)
$$\text{Al}(\text{OH})_3 + \text{NaOH} \rightarrow \text{Na}[\text{Al}(\text{OH})_4]$$

In both cases, the reaction products are salts, but in the reaction with bases they are complexes where the anion also includes a metal that originates from the hydroxide.

312. What volume of 0.50 M KOH would be required to neutralize completely 500 mL of 0.25 M H_3PO_4 solution?
A. 2.5×10^2 mL B. 1.4×10^3 mL C. 83 mL
D. 7.5×10^2 mL E. 5.2×10^2 mL

Brief:

- ❖ $\text{H}_3\text{PO}_4(\text{aq}) + 3\text{KOH}(\text{aq}) \rightarrow \text{K}_3\text{PO}_4(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$.
- Mole of H_3PO_4 = Molarity (mol/L) $\times$ mole (mol)
- Mole of H_3PO_4 = $0.25\text{mol/L} \times 0.5\text{L} = 0.125\text{mol}$ of H_3PO_4 Consuming.
- We find molar quantity of KOH (aq), using molar ratio.
- So, mole of KOH needed = $0.125 \text{ mol of } \text{H}_3\text{PO}_4 \times 3 \text{ mole of } \frac{\text{KOH}}{1 \text{ mole}} = 0.375\text{mol}$ of KOH.
- Given the molarity of KOH (=0.5M).
- Volume of KOH = $n/M = 0.375\text{mol}/0.5\text{mol/L} = 0.75\text{L} = 750\text{ml}$.

313. A 0.6745 gram sample of acid reacts with 41.75 mL of KOH solution for complete neutralization. What is the molarity of the KOH solution? (Molecular weight of Acid = 204 g/mol)
A. 0.158 M B. 0.099 M C. 0.139 M D. 0.079 M E. 0.061 M

Brief:

➤ **Given**

- ✚ The mass of Acid (HA) = 0.6745 gram
- ✚ Volume of KOH = 41.75 mL = 0.04175 L
- ✚ Molecular weight of Acid (HA) = 204 g/mol

➤ **Required**

- ✚ Molarity of KOH = ?

➤ **Solution:**

- ✚ Mole of HA = Given mass / Molecular mass
- ✚ Mole of HA = 0.6745 gram / 204 g/mol = 0.003 mol
- ✚ Knowing that the reaction between HA and KOH is a 1:1 reaction, the moles of KOH is also 0.003 mol. Molarity (M) is defined as moles of solute per liter of solution. Converting 41.75 mL to liters gives 0.04175 L.

Therefore, the molarity of the KOH solution is 0.003 mol / 0.04175 L = **0.079 M**.

314. How many equivalents of phosphoric acid are contained in 300 mL of 4.00 M phosphoric acid? (Assume the acid is to be completely neutralized by a base.)

- A. 0.6 eq B. 1.2 eq C. 2.4 eq D. 3.60 eq

- ✓ Normality = No of Eq/V(L) of solution
- ✓ No of Eq = Normality (eq/L) x V(L)
- ✓ (N = M x n) = 4 mol/L x 3 eq/mol = 12 eq/L
- ✓ No of Eq = 12 eq/L x 0.3 L = 3.6 equivalents.

315. Calculate the normality of a solution that contains 4.5 g of (COOH)₂ in 3000 mL of solution? (Assume the (COOH)₂ is to be completely neutralized in an acid-base reaction.)

- A. 0.033 N B. 0.045 N C. 0.066 N D. 0.090 N E. 0.12 N

Brief:

Normality in Chemistry is one of the expressions used to measure the concentration of a solution. It is abbreviated as **N** and is sometimes referred to as the equivalent concentration of a solution. It is mainly used as a measure of reactive species in a solution and during titration reactions or particularly in situations involving acid-base chemistry.

$$\text{Normality (N)} = \frac{\text{No of equivalent}}{\text{Volume (L) of solution}}$$
$$\text{No of equivalent} = \frac{\text{equivalent mass of solute}}{\text{molar mass of solute}}$$

$$\text{Equivalent mass} = \frac{\text{Molar mass}}{n}$$

Where n = Number of H⁺/OH⁻ furnished For acid base reaction

Number of electron transfer For Redox reaction

Magnitude of positive charge For salt compound

316. Which salt is not derived from a strong acid and a strong soluble base?

- A. MgCl_2 B. $\text{Ba}(\text{NO}_3)_2$ C. LiClO_4 D. CsBr E. NaI

Brief:

Salt	Acid	Base	Remark
MgCl_2	HCl (strong Acid)	$\text{Mg}(\text{OH})_2$ (Insoluble)	$\text{Mg}(\text{OH})_2$ is known to be insoluble in water and stable at temperatures up to 300 °C
$\text{Ba}(\text{NO}_3)_2$	HNO_3 (strong Acid)	$\text{Ba}(\text{OH})_2$ (Soluble)	
LiClO_4	HClO_4 (strong Acid)	LiOH (Soluble)	
CsBr	HBr (strong Acid)	CsOH (Soluble)	
NaI	HI (strong Acid)	NaOH (Soluble)	

317. In a sample of pure water, only one of the following statements is always true at all conditions of temperature and pressure. Which one is always true?

- A. $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-7} \text{ M}$ B. $[\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$
 C. $\text{pH} = \text{pOH} = 7.0$ D. $[\text{H}_3\text{O}^+] = [\text{OH}^-]$

Brief:

In sample of any solution the following are always fact

Acidic sample	Basic sample	Neutral sample
$[\text{H}_3\text{O}^+] > [\text{OH}^-]$	$[\text{H}_3\text{O}^+] < [\text{OH}^-]$	$[\text{H}_3\text{O}^+] = [\text{OH}^-]$ At any Temperature
$[\text{H}_3\text{O}^+] > 1 \times 10^{-7}$	$[\text{H}_3\text{O}^+] < 1 \times 10^{-7}$	$[\text{H}_3\text{O}^+] = 1 \times 10^{-7}$ At 25 ⁰ c only
$[\text{OH}^-] < 1 \times 10^{-7}$	$[\text{OH}^-] > 1 \times 10^{-7}$	$[\text{OH}^-] = 1 \times 10^{-7}$ At 25 ⁰ c only
$\text{pH} < \text{pOH}$	$\text{pH} > \text{pOH}$	$\text{pH} = \text{pOH}$ At any Temperature
$\text{pH} < 7$	$\text{pH} > 7$	$\text{pH} = 7$ At 25 ⁰ c only
$\text{pOH} > 7$	$\text{pOH} < 7$	$\text{pOH} = 7$ At 25 ⁰ c only

318. If K_w is 2.9×10^{-15} at 10°C, what is the pH of pure water at 10°C?

- A. 6.72 B. 7.00 C. 7.27 D. 7.53

Brief:

- ✓ $[\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14}$ at 25°C
- ✓ As Temperature of solution increased, the equilibrium constant of (K_w) of water is also increase and vice versa.
- ✓ $[\text{H}_3\text{O}^+][\text{OH}^-] = 2.9 \times 10^{-15}$ at 10°C
- ✓ For Pure water: $[\text{H}_3\text{O}^+] = [\text{OH}^-]$
- ✓ After simplifying $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 0.53 \times 10^{-7} = 5.3 \times 10^{-8}$
- ✓ $\text{pH} = -\log([\text{H}_3\text{O}^+]) = -\log(5.3 \times 10^{-8}) = 7.27$

319. The pOH of a solution of NaOH is 11.30. What is the $[\text{H}^+]$ for this solution?

- A. 2.0×10^{-3} B. 2.5×10^{-3} C. 5.0×10^{-12} D. 4.0×10^{-12} E. 6.2×10^{-8}

Brief:

- ❖ Given $\text{pOH} = 11.3$
- ❖ $[\text{OH}^-] = \text{antilog of pOH} = 10^{-\text{pOH}} = 10^{-11.3} = 1 \times 10^{-11.3}$
- ❖ $[\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14}$
- ❖ $[\text{H}_3\text{O}^+] = 1 \times 10^{-14} / [\text{OH}^-] = 1 \times 10^{-2.7}$
- ❖ $\text{pH} = -\log(1 \times 10^{-2.7}) = 2.0 \times 10^{-3}$

320. The $[\text{H}_3\text{O}^+]$ in a 0.050 M solution of $\text{Ca}(\text{OH})_2$ is:

- A. $1.0 \times 10^{-5} \text{ M}$ B. $5.0 \times 10^{-2} \text{ M}$ C. $1.0 \times 10^{-13} \text{ M}$ D. $5.0 \times 10^{-10} \text{ M}$ E. $2.0 \times 10^{-5} \text{ M}$

Brief:

- $\text{Ca}(\text{OH})_2 (\text{aq}) \rightarrow \text{Ca}^{2+} + 2\text{OH}^-$
 $0.05\text{M} \quad \quad 0.05\text{M} \quad 2(0.05\text{M})$
- From the dissociation equation; There is 0.1M of OH^-
- $[\text{H}_3\text{O}^+] = 1 \times 10^{-13}$
- $\text{pH} = -\log(1 \times 10^{-13}) = 13$
- $\text{pOH} = 14 - 13 = 1$.
- Solution is too much basic nature.

321. What is the pH of 500 mL of solution containing 0.0124 grams of $\text{Ca}(\text{OH})_2$?

- A. 11.04 B. 9.68 C. 2.96 D. 3.17 E. 10.83

Brief:

- ❖ The first step for this problem is calculating the concentration of $\text{Ca}(\text{OH})_2$
- ❖ $\text{Molarity} = \frac{\text{mole of } \text{Ca}(\text{OH})_2}{\text{Volume of solution}}$
- ❖ $\text{Molarity} = \frac{\text{Given mass}}{\text{molar mass} \cdot \text{Volume of solution}}$
- ❖ $\text{Molarity} = \frac{0.0124\text{gram}}{74\text{g/mol} \cdot 0.5\text{L}} = 3.3 \times 10^{-4}$
- ❖ $[\text{OH}^-] = 2(3.3 \times 10^{-4}) = 6.6 \times 10^{-4}$
- ❖ $\text{pOH} = 3.2 \quad \text{pH} = 10.8$

322. The pH of a 0.02 M solution of an unknown weak acid is 3.7. What is the pK_a of this acid?

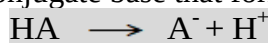
- A. 5.7 B. 4.9 C. 3.2 D. 2.8 E. 3.7

Brief:

- ❖ Knowing the pH, you know the concentration of hydrogen
 $-\log[\text{H}^+] = \text{pH} = 3.7$

$$[\text{H}^+] = 10^{-3.7} \text{ M}$$

- ❖ Now, since the weak (monoprotic) acid dissociates into its conjugate base and a proton, the mols of protons are equimolar with the mols of conjugate base the protons came FROM the weak acid, so the conjugate base that forms *must* be equimolar with the protons given out to the solvent.



$$K_a = \frac{[\text{H}^+]_{\text{eq}}^2}{[\text{HA}]_{\text{eq}}}$$

Hence, $[\text{A}^-] = [\text{H}^+]$ in the same solution volume. Using the equilibrium constant expression, we get:

$= \frac{[\text{H}^+]_{\text{eq}}^2}{[\text{HA}]_i - [\text{H}^+]_{\text{eq}}}$ $= \frac{(10^{-3.7}\text{M})^2}{0.02\text{M} - 10^{-3.7}\text{M}}$	Thus: $\text{pKa} = -\log(K_a)$ $= -\log(2.01059 \times 10^{-6})$ ≈ 5.70
--	---

323. A 0.10 M solution of a weak acid, HA, is 0.059% ionized. Evaluate K_a for the acid.
 A. 3.8×10^{-9} B. 6.5×10^{-7} C. 7.0×10^{-6} D. 4.2×10^{-6} E. 3.5×10^{-8}

Brief:

- ❖ % of ionization = $\frac{\text{ionized conc of Acid/base}}{\text{Initial conc of acid/base}} \times 100\%$
- ❖ $0.059\% = \frac{\text{ionized conc of Acid/base}}{0.1M} \times 100\%$
- ❖ Since the acid is only 0.059% ionized, therefore the concentration of ions in solution
- ❖ Conc of ionized acid (A^-) = $\frac{(0.059\%)}{100\%} \times 0.1M = 5.9 \times 10^{-5}$
- $[A^-] = [H^+]$
- ❖ $K_a = \frac{[H^+][A^-]}{[HA]} = (0.000059)^2 / 0.1 = 3.5 \times 10^{-8}$

324. If the pH of aqueous solution of 0.15 M HF is 2, what is the percent ionization of HF
 A. 2.5% B. 6.6% C. 7.5% D. 8.6%

Brief:

- ❖ % of ionization = $\frac{\text{ionized conc of Acid/base}}{\text{Initial conc of acid/base}} \times 100\%$
- ❖ The percent "ionization" is the same as the percent dissociation



We can use the Henderson-Hasselbalch equation.

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

where A^- is the conjugate base (F^-) and HA is the weak acid (HF).

$$2 = 3.15 + \log \frac{[A^-]}{[HA]}$$

$$-1.15 = \log \frac{[A^-]}{[HA]}$$

$$10^{-1.15} = \frac{[A^-]}{[HA]}$$

- ❖ But we know that as HF dissociates into F^- , the same number of moles of F^- is created as the number of moles of HF dissociated, because there is no other source that gives F^- .
- ❖ Additionally, both species are within the same solution, so their concentrations are determined by the same total volume. When we compare concentrations, we get:

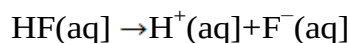
$$\frac{[A^-]}{[HA]} = \frac{\frac{\text{mol } A^-}{L_{\text{soln}}}}{\frac{\text{mol HA}}{L_{\text{soln}}}} = \frac{\text{mol } A^-}{\text{mol HA}}$$

Furthermore, we are looking for a **percent**. If we let x be the mols of F^- produced, then we can normalize the right half of the equation to say that the moles of HF leftover is $1-x$:

$$\begin{aligned} 10^{-1.15} &= \frac{x}{1-x} \\ 10^{-1.15} - 10^{-1.15}x &= x \\ 10^{-1.15} &= (1 + 10^{-1.15})x \\ x &= \frac{10^{-1.15}}{1 + 10^{-1.15}} \\ &= 0.066 \Rightarrow \mathbf{6.6\% \text{ dissociated}} \end{aligned}$$

Here's an alternative approach you can use to find the *percent ionization* of the weak acid.

- ❖ When placed in aqueous solution, *hydrofluoric acid*, HF, will dissociate to give *hydrogen cations*, H^+ , and *fluoride anions*, F^- .



- ❖ Notice that you have 1:1 mole ratios between all the species that take part in this reaction. This means that every mole of hydrofluoric acid that dissociates will produce one mole of hydrogen cations and one mole of fluoride anions.
- ❖ Simply put, one molecule of hydrofluoric acid that ionizes will split into one hydrogen cation and one fluoride anion.
- ❖ Use the pH of the solution to determine the *concentration* of hydrogen cations
 - $[H^+] = 10^{-pH}$, $[H^+] = 10^{-2} = 0.010 \text{ M}$
- ❖ So, if the ionization of the acid produced 0.010 M of H^+ , it follows that it also produced 0.010 M of F^- .
- ❖ This means that out of the initial 0.15 M of weak acid, only 0.010 M ionized.

$$\begin{aligned} \% \text{ ionization} &= \frac{\text{amount that ionized}}{\text{Initial amount}} \times 100 \\ \% \text{ ionization} &= \frac{0.01\text{M}}{0.15\text{M}} \times 100 = 6.6\% \end{aligned}$$

325. Which of the following weak acids ionizes to give the strongest conjugate base?
 A. HClO B. CH₃COOH C. HF D. HNO₂ E. HCN

Brief:

❖ According to Bronsted lowry definition

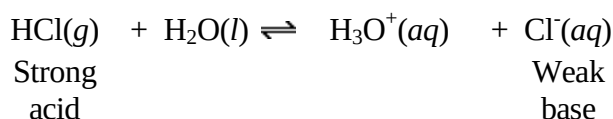
✓ Acids are proton acceptor

✓ Bases are proton donor

The Relative Strengths of Conjugate Acid-base Pairs

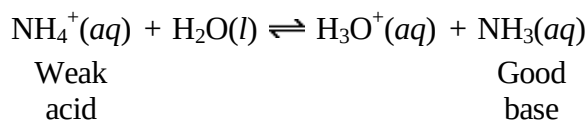
- *Strong acids* have a weak conjugate base.

Example: HCl is a strong acid. If HCl is a strong acid, it must be a good proton donor. HCl can only be a good proton donor, however, if the Cl⁻ ion is a poor proton acceptor. Thus, the Cl⁻ ion must be a weak base.



- *Strong bases* have
- a weak conjugate acid.

Example: Let's consider the relationship between the strength of the ammonium (NH₄⁺) and its conjugate base, ammonia (NH₃). The NH₄⁺ ion is a weak acid because ammonia is a reasonably good base.



Comparing Relative Strengths of Pairs of Acids and Bases

The value of K_a for an acid can be used to decide whether it is a strong acid or a weak acid, in an absolute sense. It can also be used to compare the relative strengths of a pair of acids.

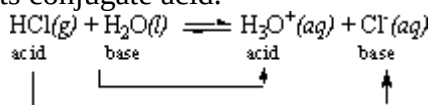
Example: Consider HCl and the H₃O⁺ ion.

$$\text{HCl} \quad K_a = 1 \times 10^6$$

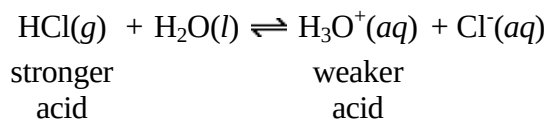
$$\text{H}_3\text{O}^+ \quad K_a = 55$$

These K_a values suggest that both are strong acids, but HCl is a stronger acid than the H₃O⁺ ion.

A high proportion of the HCl molecules in an aqueous solution reacts with water to form H₃O⁺ and Cl⁻ ions. The Brnsted theory suggests that every acid-base reaction converts an acid into its conjugate base and a base into its conjugate acid.



There are two acids and two bases in this reaction. The stronger acid, however, is on the left side of the equation.

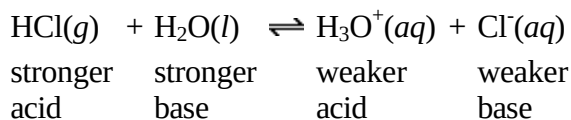


The general rules suggest that the stronger of a pair of acids must form the weaker of a pair of conjugate bases. The fact that HCl is a stronger acid than the H_3O^+ ion implies that the Cl^- ion is a weaker base than water.

Acid strength: $\text{HCl} > \text{H}_3\text{O}^+$

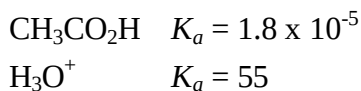
Base strength: $\text{Cl}^- < \text{H}_2\text{O}$

Thus, the equation for the reaction between HCl and water can be written as follows.

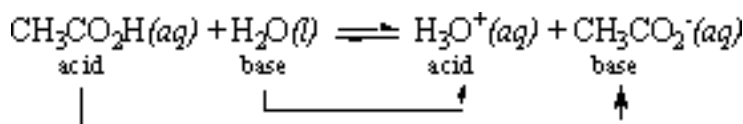


It isn't surprising that 99.996% of the HCl molecules in a 6 M solution react with water to give H_3O^+ ions and Cl^- ions. The stronger of a pair of acids should react with the stronger of a pair of bases to form a weaker acid and a weaker base.

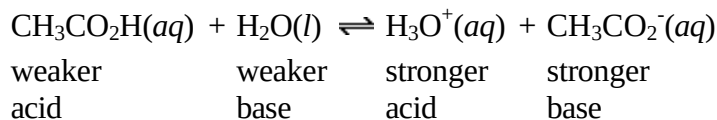
Let's look at the relative strengths of acetic acid and the H_3O^+ ion.



The values of K_a for these acids suggest that acetic acid is a much weaker acid than the H_3O^+ ion, which explains why acetic acid is a weak acid in water. Once again, the reaction between the acid and water must convert the acid into its conjugate base and the base into its conjugate acid.



But this time, the stronger acid and the stronger base are on the right side of the equation.



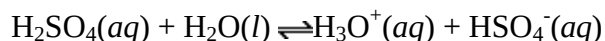
As a result, only a few of the $\text{CH}_3\text{CO}_2\text{H}$ molecules actually donate an H^+ ion to a water molecule to form the H_3O^+ and CH_3CO_2^- ions.

NaNH_2 reacts with water to give an aqueous solution of NaOH and NH_3 .



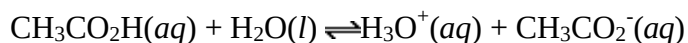
Determine which is the stronger base, the NH_2^- or OH^- ions.

The magnitude of K_a can also be used to explain why some compounds that qualify as Brnsted acids or bases don't act like acids or bases when they dissolve in water. When the value of K_a for an acid is relatively large, the acid reacts with water until essentially all of the acid molecules have been consumed. Sulfuric acid ($K_a = 1 \times 10^3$), for example, reacts with water until 99.9% of the H_2SO_4 molecules in a 1 M solution have lost a proton to form HSO_4^- ions.



As K_a becomes smaller, the extent to which the acid reacts with water decreases.

As long as K_a for the acid is significantly larger than the value of K_a for water, the acid will ionize to some extent. Acetic acid, for example, reacts to some extent with water to form H_3O^+ and CH_3CO_2^- , or acetate, ions.



As the K_a value for the acid approaches the K_a for water, the compound becomes more like water in its acidity. Although it is still a Brnsted acid, it is so weak that we may be unable to detect this acidity in aqueous solution.

Some potential Brnsted acids are so weak that their K_a values are smaller than water's. Ammonia, for example, has a K_a of only 1×10^{-33} . Although NH_3 can be a Brnsted acid, because it has the potential to act as a hydrogen-ion donor, there is no evidence of this acidity when it dissolves in water.

- ❖ HClO_4 (strong acid) ClO_4^- (weak conjugated base)
- ❖ CH_3COOH (weak acid) CH_3COO^- (strong conjugated base)
- ❖ HF (weak acid) F^- (strong conjugated base)
- ❖ HNO_2 (weak acid) NO_2^- (strong conjugated base)
- ❖ HCN (weak acid) CN^- (strong conjugated base)

326. Which of the following is true about a 0.10 M solution of a weak acid, HX?

- A. $[\text{X}^-] = 0.10 \text{ M}$ B. $\text{pH} = 1$ C. $[\text{HX}] > [\text{H}^+]$ D. $[\text{H}^+] = 0.10 \text{ M}$

Brief:

- ❖ For a weak acid HX
 - ✓ $\text{HX} + \text{H}_2\text{O} \rightleftharpoons \text{A}^- + \text{H}_3\text{O}^+$
 - ✓ Since HX is a weak acid, it dissociate slightly.
 - ✓ $[\text{X}^-] = [\text{H}^+] < 0.1\text{M}$
 - ✓ $[\text{HX}] > [\text{H}^+]$
 - ✓ $\text{pH} > 1$

327. Calculate the pH of a 0.50 M solution of NaNO_3 .

- A. 12.18 B. 5.48 C. 1.82 D. 7.00

Brief:

- ❖ Since NaNO_3 is salt of strong acid (HNO_3) and strong base (NaOH), those having weak conjugated species those are undergo neither cationic nor anionic hydrolysis reaction. As result the pH of such solution to be 7.

328. What is the pH of a 0.150 M solution of sodium acetate (NaO_2CCH_3)?
 $K_a(\text{CH}_3\text{CO}_2\text{H}) = 1.8 \times 10^{-5}$.

Brief:

- ❖ Sodium acetate is the salt of a weak acid and strong base from the equation
- ❖ $\text{C}_2\text{H}_3\text{NaO}_2 \rightarrow \text{CH}_3\text{COO}^- + \text{Na}^+$
- ❖ where: $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} ; \text{CH}_3\text{COOH} + \text{OH}^-$
- ❖ As it is a weak acid and strong base, this is a good indicator of a fairly high pH
- ❖ $K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]}$

$$\text{Also, } K_a K_b = 1 \cdot 10^{-14} / \text{So, } K_b = \frac{1 \cdot 10^{-14}}{1.8 \cdot 10^{-5}} = 5.555... \cdot 10^{-10}$$

$$\begin{array}{c} \text{CH}_3\text{COO}^- \quad \text{CH}_3\text{COOH} \quad \text{OH}^- \\ \left[\begin{array}{ccc} I & 0.150 & 0 & 0 \\ C & -x & +x & +x \\ E & 0.150 - x & x & x \end{array} \right] \end{array}$$

$$K_b = 5.555... \cdot 10^{-10} = \frac{x^2}{0.150 - x}$$

Since the value for K_b is small, we will assume the the value for x is small, and so will take 0.150.

$$\text{So, } x = \sqrt{0.150(5.555... \cdot 10^{-10})} = \text{OH}^-$$

$$\begin{aligned} \text{pH} &= 14 - \text{pOH} = 14 - (-\log(\text{OH}^-)) \\ &= 14 - (-\log(\sqrt{0.150(5.555... \cdot 10^{-10})})) \\ &= 14 - (-\log(9.128709292 \cdot 10^{-6})) \\ &= 14 - 5.039590623 \approx 8.96 \end{aligned}$$

329. Buffer solutions

- Will always have a pH of 7.
- Are rarely found in living systems.
- Cause a decrease in pH when acids are added to them.
- Tend to maintain a relatively constant pH.
- Cause an increase in pH when acids are added to them.

Brief.

Characteristics of buffer solution

- ❖ It has a definite pH.
- ❖ Its pH does not change on standing for long periods of time.
- ❖ Its pH does not change on dilution.
- ❖ Its pH is slightly changed by the addition of small quantity of an acid or base.
- ❖ Tend to maintain a relatively constant pH.
- ❖ It is able to neutralize small amounts of added acid or base.
- ❖ Maintaining the pH of the solution relatively stable.
- ❖ The pH will not change if a small amount of concentrated or strong acid or base is added.
- ❖ This is because a buffer solution consists of a conjugate acid-base pair that neutralizes the acid or base added and resists the change in the pH.
- ❖ The three primary major buffering systems in the body are:
 - ✓ Carbonic Acid-Bicarbonate Buffer.
 - ✓ Phosphate Buffer System.
 - ✓ Protein Buffer System.

330. Which sentence is false about Henderson-Hasselbalch equation ?

- A. It is used to determine pKa value of solutions.
- B. PH of solutions containing comparable and appreciable conc. of a conjugate acid-base pair to ratio of their concentrations.
- C. It is used to calculate valency of molecule.
- D. $\text{pH} = \text{pka} + \log [\text{conjugated base}]/[\text{conjugated acid}]$

Brief:

➤ Facts about Henderson-Hasselbalch equation

- ❖ The Henderson-Hasselbalch equation gives the reader an interaction between the acid dissociation constant (pKa) and the pH of acids (in aqueous solutions).
- ❖ The Henderson equation is a chemical equation that relates the acid dissociation constant to the pH of the aqueous solution of an acid.
- ❖ The Henderson-Hasselbalch equation is used under the condition of weak acid and a salt or a weak base and a salt, both of these are buffer solutions.
- ❖ The four variables present in the equation: pH, pKa, $[A^-]$ and $[HA]$.

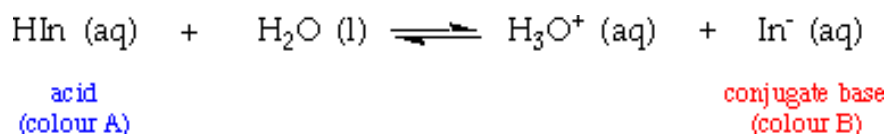
- ❖ If the pH of the solution is known as is the pKa for the acid, the ratio of conjugate base to acid can be determined
- ❖ At pH = pKa for an acid, there is an equal amount of conjugate base to acid
- ❖ At pH > pKa for an acid, the acid will be mostly ionized.
- ❖ At pH < pKa for an acid, the acid will be mostly non-ionized.
- ❖ The "operational" relationship between pKa and pH is mathematically represented by Henderson-Hasselbach equation: $\text{pH} = \text{pK}_a + \log [\text{A}^-] / [\text{HA}]$

331. Which sentence is false about indicator?

- A. Phenolphthalein gives colorless solutions at pH 8.3-11.0 in acidic medium.
- B. Methyl red gives red solutions at pH 4.2-6.6 in acidic medium.
- C. Methyl orange give orange solutions at pH 3.1-4.4 in base medium.
- D. Methyl red gives red solutions at pH 4.2-6.6 in basic medium.

Brief

- ❖ Titration: pH indicator is an auxiliary substance added for physical detection of completion of a titration at its equivalence point. It generally shows colour change on the completion of the titration.
- ❖ Indicators : Indicators are substances which indicate the end-point of a titration by changing their colour.
- ❖ Indicators are either weak organic acids or weak organic bases having characteristically different colors in ionized and unionized forms.
- ❖ Acid - Base indicators (also known as pH indicators) are substances which change colour with pH. They are usually weak acids or bases, which when dissolved in water dissociate slightly and form ions.
- ❖ Consider an indicator which is a weak acid, with the formula HIn.



- ❖ An indicator's color is the visible result of the ratio of the concentrations of the two species In^- and HIn. If most of the indicator (typically about 60–90% or more) is present as In^- , then we see the color of the In^- ion, which would be yellow for methyl orange. If most is present as HIn, then we see the color of the HIn molecule:

$$\frac{[\text{In}^-]}{[\text{HIn}]} = \frac{[\text{substance with yellow color}]}{[\text{substance with red color}]} = \frac{K_a}{[\text{H}_3\text{O}^+]}$$

- ❖ This shows us how the ratio of $[\text{In}^-]/[\text{HIn}]$ varies with the concentration of hydronium ion. The above expression describing the indicator equilibrium can be rearranged.

$$\begin{aligned}\frac{[\text{H}_3\text{O}^+]}{K_a} &= \frac{[\text{HIn}]}{[\text{In}^-]} \\ \log\left(\frac{[\text{H}_3\text{O}^+]}{K_a}\right) &= \log\left(\frac{[\text{HIn}]}{[\text{In}^-]}\right) \\ \log([\text{H}_3\text{O}^+]) - \log(K_a) &= -\log\left(\frac{[\text{In}^-]}{[\text{HIn}]}\right) \\ -\text{pH} + \text{p}K_a &= -\log\left(\frac{[\text{In}^-]}{[\text{HIn}]}\right) \\ \text{pH} = \text{p}K_a + \log\left(\frac{[\text{In}^-]}{[\text{HIn}]}\right) &\text{ or } \text{pH} = \text{p}K_a + \log\left(\frac{[\text{base}]}{[\text{acid}]}\right)\end{aligned}$$

Indicator	Colour		$\text{p}K_{\text{In}}$	pH range
	Acid	Base		
Thymol Blue	red	yellow	1.5	1.2 - 2.8
Methyl Orange	red	yellow	3.7	3.2 - 4.4
Bromocresol Green	yellow	blue	4.7	3.8 - 5.4
Methyl Red	yellow	red	5.1	4.8 - 6.0
Bromothymol Blue	yellow	blue	7.0	6.0 - 7.6
Phenol Red	yellow	red	7.9	6.8 - 8.4
Phenolphthalein	colorless	pink	9.4	8.2 - 10.0
Litmus	Red	blue	6.3	5.0 – 8.0
Methyl Violet	Yellow	Violet	9.4	0.0 - 1.6

332. What is an Equivalence point?

- A. Amount of titrant = Amount of analyte in a sample.
- B. Amount of titrant > Amount of analyte in a sample.
- C. Amount of titrant < Amount of analyte in a sample.
- D. None of the above

Brief:

❖ Equivalence point:

- ✓ The volume of added base required to exactly satisfy the reaction stoichiometry.
- ✓ A point in titration at which the amount of titrant added is just enough to completely neutralize the analyte solution.
- ✓ At the equivalence point in an acid-base titration, moles of base = moles of acid and the solution only contains salt and water.
- ✓ The pH at the equivalence point = 7 for a strong acid/strong base titration.
- ✓ The pH at the equivalence point >7 for a weak acid/strong base titration.
- ✓ The pH at the equivalence point < 7 for a weak base/strong acid titration.

333. What is the end point?

- A. Physical change occurs which related to the chemical change.
- B. A reagent of a known concentration which used in the titrimetric analysis.
- C. Solutions of one reactant of accurately known concentration.
- D. Amount of titrant = Amount of analyte in a sample.

Brief:

❖ The endpoint of the titration is

- ✓ The point at which the colour of the solution changes completely due to the formation of product due to the addition of indicator.
- ✓ It is the point in a titration when a physical change occurs that is associated with the condition of chemical equivalence.
- ✓ It is a point in a titration that signifies the completion of the titration by a change in the colour or intensity of the solution.
- ✓ Indicator colour change is the end point of the titration
- ✓ The end point is used as an approximation of the equivalence point and is employed, with the known concentration of the titrant, to calculate the amount or concentration of the analyte.

- ✓ If a few drops of indicator solution have been added, this sharp increase in pH causes an abrupt change in colour, which is called the endpoint of the indicator. The terms endpoint and equivalence point are not interchangeable. The titration is over when the endpoint is reached and the indicator changes colour.

Categories	Endpoint	Equivalence point
Definition	The point at which color change occurs due to a pH-change	A point when the chemical reaction in the titration mixture ends
Explanation	At this point, the moles of titrant exceed the moles of the analyte.	A point in a titration when the number of moles of titrant are equal to analyte
Occurrence	After equivalence point	Before the endpoint
Completion	Titration is complete once the endpoint reaches	It is not the completion of a titration reaction
Reaction completion	It does not mean the completion of reaction between analyte and titrant	It means the completion of reaction between titrant and analyte
Times of occurrence	Occurs only once in a reaction	Titration processes can have multiple equivalence points
Nutshell	Change in color	Just before change in color

334. Which of the following is the correct pH value during a titration when 9.54 mL of a 0.15 M HCl solution has reacted with 22.88 mL of 0.14 M NaOH?
- A. 1.25 B. 6.22 C. 12.75 D. 8.87

Brief:

- ❖ Calculating the number of moles of hydrochloric acid and the number of moles of sodium hydroxide that took part in the reaction.

$$9.54 \text{ mL solution} \cdot \frac{0.15 \text{ moles HCl}}{10^3 \text{ mL solution}} = 0.001431 \text{ moles HCl}$$

$$22.88 \text{ mL solution} \cdot \frac{0.14 \text{ moles NaOH}}{10^3 \text{ mL solution}} = 0.003203 \text{ moles NaOH}$$

- ❖ Based on the reaction given below, hydrochloric acid and sodium hydroxide neutralize each other in a 1:1 mole ratio. $\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$
- In order to have a complete neutralization, it need to mix equal numbers of moles of hydrochloric acid and of sodium hydroxide

- ❖ In this case there is more moles of sodium hydroxide than of hydrochloric acid, so that the acid will act as a limiting reagent, i.e. it will be completely consumed *before* all the moles of sodium hydroxide will get the chance to react.

This means that after the reaction takes place, solution will contain 0 moles of hydrochloric acid and $0.003 \text{ moles} - 0.0014 \text{ moles} = 0.00178 \text{ moles NaOH}$

- ❖ At this point, the pH of the resulting solution must be >7 because the left solution that contains a strong base.
- ❖ Since sodium hydroxide dissociates completely in aqueous solution to produce sodium cations and hydroxide anions: $\text{NaOH(aq)} \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$

This means that the resulting solution will contain 0.00178 moles of hydroxide anions.

- ❖ The total volume of the resulting solution will be: $9.54 \text{ mL} + 22.88 \text{ mL} = 32.42 \text{ mL}$
- ❖ The molarity of the hydroxide anions will be

$$[\text{OH}^-] = 0.0018 \text{ moles} / 3.2 \times 10^{-3} \text{ L} = 0.055 \text{ mol L}^{-1}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log (0.055) = 1.25$$

$$\text{pH} = 14 - \text{pOH} = 14 - 1.25 = 12.75$$

335. How many moles of Unknown weak acid (HA) in a 100L solution are required to produce a solution with $\text{pH}=2$? K_a of the Unknown weak acid (HA) = 5×10^{-3} (self-practice)

A. 3×10^{-3} B. 0.3×10^{-3} C. 3 D. 0.3

336. How can a 100mL of sodium hydroxide solution with a pH of 13.00 be converted to a sodium hydroxide solution with a pH of 12.00?

A. By diluting a solution with distilled water to a total volume of 100 ml
 B. By diluting a solution with distilled water to a total volume of 200 ml
 C. By diluting a solution with distilled water to a total volume of 1.00 L
 D. By adding 100 ml of NaOH 0.10 M NaOH

Brief:

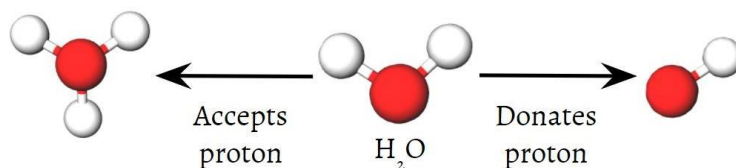
- If the pH of a NaOH solution = 13, then $\text{pOH} = (14 - 13) = 1$
 So $[\text{OH}^-] = 1 \times 10^{-1} = 0.1 \text{ M}$
- If the pH of the NaOH solution is 12, then $\text{pOH} = 2$
 So it's $[\text{OH}^-] = 1 \times 10^{-2} = 0.01$
- Therefore to decrease the pH of NaOH solution from 13 to 12 its concentration is to be decreased 10 times.

The concentration of OH^- by 10 times when the volume of solution increased from 100ml to 1000ml (by adding of 900ml) by applying the dilution formula: $M_i V_i = M_f V_f$

- iv. So converting 100 ml of 0.1M NaOH to 1000 ml of 0.01M NaOH, the pH can be reduced to 12.
- v. So if we add 900 ml of distilled water to 100 ml of 0.1M NaOH solution (of pH= 13) we will get 1000 ml of 0.01M NaOH , (pH will be = 12).
337. Which of the following is not amphoteric species?
- A. H_2O B. H_2SO_4 C. HCO_3^- D. H_2PO_4^-

Brief:

- ❖ Molecules or ions which can either donate or accept a proton, depending on their circumstances, are called amphoteric species.
- ❖ The most important amphoteric species is water itself. When an acid donates a proton to water, the water molecule is a proton acceptor, and hence a base.



- ❖ Conversely, when a base reacts with water, a water molecule donates a proton, and hence acts as an acid.
 - ❖ Other Examples: amino acids, hydrogencarbonate ion (or bicarbonate ion) HCO_3^- , dihydrogen phosphate ion H_2PO_4^- , and hydrogensulfate ion (or bisulfate ion) HSO_4^-
338. Which of the following acid is the strongest?
- A. HCN , $K_a = 7.2 \times 10^{-10}$ C. HCOOH , $k_a = 1.8 \times 10^{-4}$
 B. CH_3COOH , $k_a = 1.8 \times 10^{-5}$ D. $\text{C}_6\text{H}_5\text{COOH}$, $K_a = 6.3 \times 10^{-3}$

Brief:

- ❖ The acid dissociation constant (K_a) is used to differentiate between strong and weak acids. The acid dissociates more as the K_a increases. Strong acids must therefore dissociate more in water. A weak acid, on the other hand, is less likely to ionise and release a hydrogen ion, leading to a less acidic solution.
- ❖ The equilibrium constant is a numerical representation of an acid's strength in a solution. K_a is often stated in mol/L units.
- ❖ K_a can be used to calculate the equilibrium position:
 - When K_a is big, the creation of dissociation products is encouraged.
 - When K_a is low, the undissolved acid takes precedence.
- ❖ K_a can be used to estimate an acid's strength:
 - If K_a is high (and $\text{p}K_a$ is low), the acid is largely dissociated and therefore powerful.

- When K_a is low (and pK_a is high), there has been little dissociation, hence the acid is weak.
- Because adding water to an acid solution does not change its acid equilibrium constant, but it does modify the H^+ ion concentration and pH, K_a is a better estimate of an acid's strength than pH.

339. A buffer solution cannot be prepared from a mixture of :-

- A. $HCOOH$ and $HCOOK$ C. NH_3 and NH_4Cl
 B. CH_3COOH and CH_3COOK D. KOH and K_2SO_4

Brief:

- ❖ Buffer solution is not prepared from strong acid and strong base, this is because they ionize completely, like KOH and H_2SO_4

340. Which acid-base indicator will give the best results for the titration of an aqueous ammonia solution with 0.10 M HCl ?

- A. Methyl violet B. Cresol red C. Methyl red D. Phenolphthalein

Indicator	Color change	pH range
Methyl Violet	Red → blue	0 - 2
Methyl Red	Red → yellow	4 - 6
Cresol Red	Yellow → purple	7 - 9
Phenolphthalein	Colorless → pink	8 - 10

UNIT EIGHT

8. ELECTROCHEMISTRY

Unit out come

- ***After completing this unit, you will be able to:***
- ❖ *Describe the fundamental concepts related to oxidation- reduction reaction*
 - ❖ *Developing the skill of determining the oxidation number of elements*
 - ❖ *Distinguish between electrolytic cell and galvanic cell*
 - ❖ *Explain the application of redox reactions in production of new substances and energy*
 - ❖ *Explain the inter conversion of chemical energy and electrical energy*
 - ❖ *Distinguish between metallic conduction and electrolytic conduction*
 - ❖ *Identify and describe the functioning of the components of electrolytic and galvanic cells*
 - ❖ *Distinguish difference between electrolytic and galvanic cells*
 - ❖ *Developing skill of problem solving by Applying of Faradays laws to solve problems associated with electrolysis*
 - ❖ *Explain the manufacturing of metals, non- metals and chemicals through electrolysis process*
 - ❖ *Developing skills of identifying whether cell reaction is occur spontaneously or non-spontaneously.*
 - ❖ *Calculate the emf of an electrochemical cell experimentally or from given data*

8.1. Introduction

Electrochemistry is the branch of chemistry that studies the relationship between electricity and chemical reactions. It deals with the movement of electrons in redox reactions, which are essential for batteries, corrosion, electrolysis, and fuel cells.

8.1.1. Redox Reactions and Electrochemical Cells

- ✓ Oxidation: Loss of electrons
- ✓ Reduction: Gain of electrons
- ✓ Oxidizing agent: Accepts electrons (gets reduced)
- ✓ Reducing agent: Donates electrons (gets oxidized)

❖ Types of Electrochemical Cells

Galvanic (Voltaic) Cell – Converts chemical energy into electrical energy (e.g., batteries).

Electrolytic Cell – Uses electrical energy to drive a non-spontaneous reaction (e.g., electrolysis).

Galvanic Cell Components

1. Anode (-ve electrode) → Where oxidation occurs.
2. Cathode (+ve electrode) → Where reduction occurs.
3. Salt Bridge → Maintains charge balance by allowing ion flow.
4. Electrolyte → Conducting solution containing ions.

Example: Daniel Cell (Zn-Cu Cell)

- * $\text{Zn(s)} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ Oxidation at anode
- * $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu(s)}$ Reduction at cathode)

8.1.2. Standard Electrode Potential (E°)

1. The tendency of a half-cell to gain or lose electrons under standard conditions (1M, 1 atm, 25°C).
2. Measured in volts (V) using the Standard Hydrogen Electrode (SHE) as reference.
3. Electrochemical Series: Lists electrodes in order of increasing E° values.
 - Cell Potential Calculation
 - $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$ If $E^\circ_{\text{cell}} > 0$, the reaction is spontaneous.
 - Nernst Equation is used to calculate electrode potential at non-standard conditions:

$$E = E^\circ - \frac{0.0591}{n} \log Q$$

where Q is the reaction quotient, and n is the number of electrons transferred.

➤ Electrolysis

- Process of using electricity to drive a chemical reaction.
- Occurs in electrolytic cells (non-spontaneous reactions).

❖ Faraday's Laws of Electrolysis

- ✓ First Law: The mass of a substance deposited (m) is directly proportional to the charge (Q).
 $m = Z \cdot Q$
- ✓ Second Law: The masses of different substances deposited by the same charge are proportional to their equivalent weights.

8.1.3. Batteries and Fuel Cells

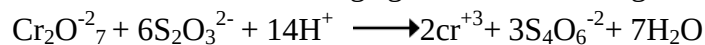
- ✓ Primary Batteries (non-rechargeable) – e.g., Dry cell, Alkaline battery.
- ✓ Secondary Batteries (rechargeable) – e.g., Lead-acid battery, Lithium-ion battery.
- ✓ Convert chemical energy directly into electrical energy.
- ✓ Example: Hydrogen Fuel Cell
 - Anode: $\text{H}_2 \rightarrow 2\text{H} + 2\text{e}^-$
 - Cathode: $\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$
 - Overall reaction: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{Energy}$

8.1.4. Corrosion and Its Prevention

- ✓ Corrosion: Undesired oxidation of metals (e.g., rusting of iron).
- ✓ Methods of Prevention:
 - Galvanization (coating with zinc)
 - Cathodic Protection (sacrificial anode)
 - Alloying (stainless steel)
 - Painting or coating

8.2. Questions with Explanation

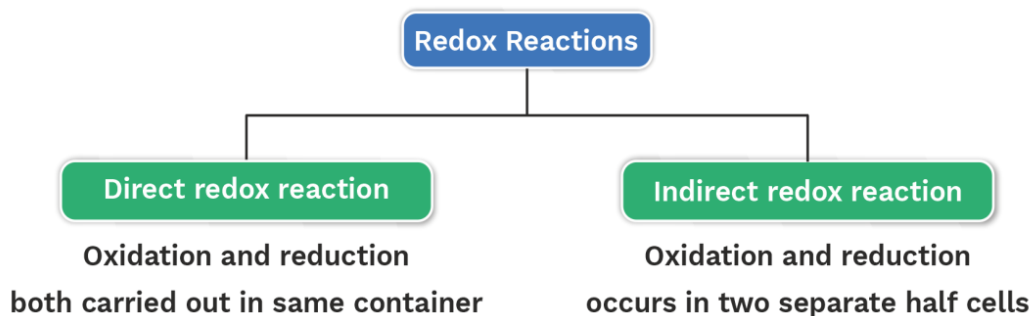
341. Which substance is reducing agent in the following reaction?



A. $\text{S}_2\text{O}_3^{2-}$ B. Cr^{+3} C. S D. H^+

Brief:

- ❖ Calculate the oxidation number of each atoms
- ❖ Identify under which atoms change in oxidation number
- ❖ Atom with increasing oxidation number is oxidized (reducing agent)
- ❖ Atom with decreasing oxidation number is reduced (oxidizing agent)
- ❖ The oxidation number of Cr decreased from +6 to +3.....oxidation
- ❖ The oxidation number of S increased from +2 to +2.5...reduction
- ❖ $\text{Cr}_2\text{O}_7^{2-} + 6\text{S}_2\text{O}_3^{2-} + 14\text{H}^+ \longrightarrow 2\text{Cr}^{+3} + 3\text{S}_4\text{O}_6^{2-} + 7\text{H}_2\text{O}$



Oxidation and reduction

	Property	Oxidation	Reduction
1	Hydrogen	Loss	Gain
2	Oxygen	Gain	Loss
3	Electron	Loss	Gain
4	Oxidation No.	Increase	Decrease

342. The standard emf for the cell reaction $\text{Zn} + \text{Cu}^{+2} \longrightarrow \text{Cu} + \text{Zn}^{+2}$ is 1.1V.

what would the emf the cell reaction be when 0.1M Cu^{+2} and 0.1M Zn^{+2} solutions are used at 298K?

A. 1.1V

V. 0.10V

C. -1.1V

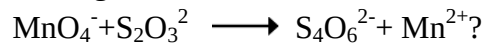
D. -0.11V

$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{0.0592}{n} \log \frac{[\text{Zn}^{+2}]}{[\text{Cu}^{+2}]}$$

Brief:

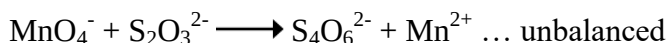
$$= 1.1\text{V} - \frac{0.0592}{2} \log \frac{[0.1\text{M}]}{[0.1\text{M}]} = 1.1\text{V}$$

343. When the equation for the following reaction in basic solution is balanced, what is the sum of the coefficients?



A. 39 B. 43 C. 23 D. 18

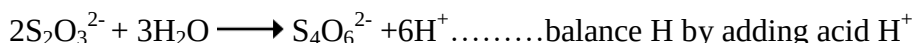
Brief:



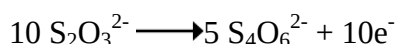
Mn goes from +7 on the left to +2 on right, so it is reduced.

S goes from +2 on the left to +2.5 on the right, so it is oxidized.

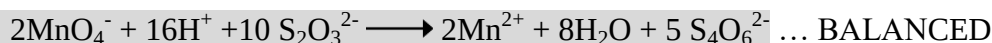
Write the oxidation and reduction cell reaction separately



To equalize electrons, multiply reduction rx by 2 and oxidation rxn by 5 to get...

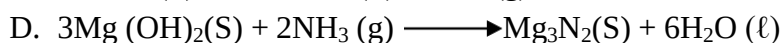
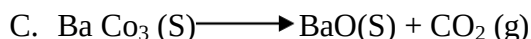
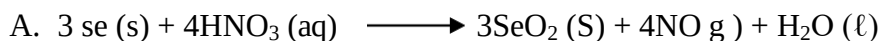


Add them together and combine/cancel like terms to obtain...



Now, you can obtain the sum of coefficients of this balanced chemical equation

344. Which of the following reactions is redox reaction



Brief:

- ❖ When a change in oxidation number occurs in a reaction, with both an increase and a decrease in oxidation number, then the reaction is classified as redox. If this does not occur, then the reaction is non-redox.

❖ An oxidation-reduction (redox) reaction is a type of chemical reaction that involves a transfer of electrons between two species. An oxidation-reduction reaction is any chemical reaction in which the oxidation number of a molecule, atom, or ion changes by gaining or losing an electron.

❖ **Common redox reaction are**

- ✓ Combination reaction.
- ✓ Decomposition reaction.
- ✓ Single displacement reaction (Metal or Non-metal displacement)
- ✓ Disproportionation reaction and Comproportionation reaction are special type of redox reaction

✚ **Disproportionation reaction**, is basically a type of redox reaction involving simultaneous reduction and oxidation of atoms of the same element from one oxidation state (OS) to two different oxidation states. Eg: $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$.

✚ **Comproportionation or synproportionation** is a chemical reaction where two reactants containing the same element but with different oxidation numbers, form a compound having an intermediate oxidation number. It is the opposite of disproportionation. Eg. $2\text{H}_2\text{S} + \text{SO}_2 \rightarrow 3\text{S} + 2\text{H}_2\text{O}$

345. An electrochemical cell constructed for the reaction

$\text{Cu}^{+2}(\text{aq}) + \text{M}(\text{S}) \longrightarrow \text{Cu}(\text{s}) + \text{M}^{2+}(\text{aq})$ has an $E_{\text{cell}} = 0.75\text{V}$. The standard reduction E^0 red potential for $\text{Cu}^{+2}(\text{aq})$ is 0.34V . What is the standard reduction potential for $\text{M}^{+2}(\text{aq})$?

- A. -1.09v B. -0.41v C. 0.41v D. 1.09v

Brief:

Use the formula: $E = E^{\circ} - \frac{0.0592 \text{ V}}{n} \log Q$

- ✓ $Q = [\text{M}^{2+}] / [\text{Cu}^{+2}]$, $n = 2$, $E_{\text{cell}} = 0.75\text{v}$, $E^0(\text{Cu}^{+2}) = 0.34\text{V}$
- ✓ $E^0_{\text{cell}} = E^0(\text{reduction at cathode}) - E^0(\text{reduction potential at Anode})$
- ✓ $E^0_{\text{cell}} = E^0(\text{reduction at cathode}) + E^0(\text{Oxidation potential at Anode})$
- ✓ Then calculate the potential of $[\text{M}^{2+}]$ and identify what metal to be M

346. What amount of electric current is needed to deposit 0.52g of chromium metal from a solution of Cr^{3+} in a period of 965 sec ?

- A. 0.408A B. 3A C. 7.9A D. 0.79A

Brief: $\text{Cr}^{3+} + 3\text{e}^- \longrightarrow \text{Cr}$

*The mass of Cr deposited on 3mole of electron and passing 96500C charge is 52 g/mol.

* 3 mole of electron = 1mole of electron = 96500C

* For Cr, There are 3mole of electron, the charge passing on the cell

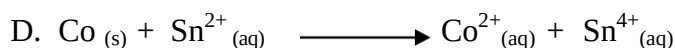
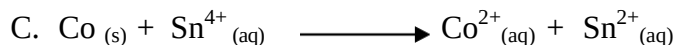
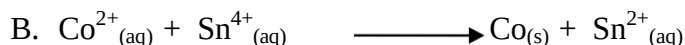
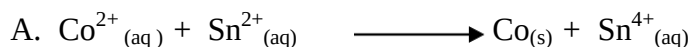
$$= 3 \times 96500 \text{C} =$$

* In 0.52g of Cr, How much the charge passing through the cell = ? 289500

* i.e $(0.52 \text{g} \times 289500 \text{C/mol}) / 52 \text{g/mol} = 2895 \text{C}$

* $Q = It$, $I = Q/t = 2895 \text{C} / 965 \text{sec} = 3 \text{A}$

347. Which one of the following reaction represents the following cell notation?



Brief:

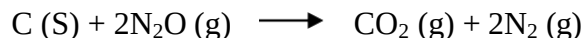


Given from the cell notation:

✓ Cell notation = Anode electrode/Anode electrolyte//Cathode electrolyte/cathode electrode, $\text{Co}_{(s)} / \text{Sn}^{4+}_{(aq)} // \text{Co}^{2+}_{(aq)} / \text{Sn}^{2+}_{(aq)}$

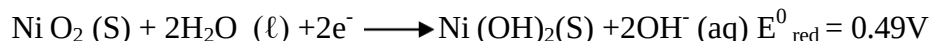
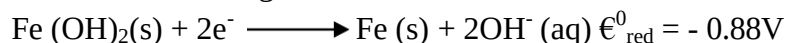
✓ $\text{Co}_{(s)} + \text{Sn}^{4+}_{(aq)} \longrightarrow \text{Co}^{2+}_{(aq)} + \text{Sn}^{2+}_{(aq)}$

348. Which substance is reduced in the reaction (self-practice)



A. C B. N_2O C. CO_2 D. N_2

349. Consider the following half – reactions



350. What is the standard cell potential for a voltaic cell using these electrode (self-practice)

A. -1.37V B. -0.39V C. 0.49V D. 1.37V

351. In chemical equilibrium what will happen when reactants are favored over product?

A. $\Delta G^\circ > 0$ & $K < 1$

C. $\Delta G^\circ > 0$ & $K > 1$

B. $\Delta G^\circ < 0$ & $K < 1$

D. $\Delta G^\circ < 0$ & $K > 1$

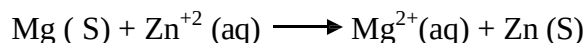
Brief:

✓ $\Delta G < 0$, $E_{\text{cell}} > 0$, $K > 1$: reaction can spontaneously proceed to the right: $A \rightarrow B$

✓ $\Delta G > 0$, $E_{\text{cell}} < 0$, $K < 1$: reaction can spontaneously proceed to the left: $A \leftarrow B$

✓ $\Delta G = 0$, $E_{\text{cell}} = 0$, $K = 1$: the reaction is at equilibrium; the quantities of [A] and [B] will not change

352. Given the following reaction



The correct cell notation for the reaction is (self- practice)

- i. $\text{Zn}^{2+}(\text{aq}) / \text{Zn(S)} // \text{Mg (S)} / \text{Mg}^{2+}(\text{aq})$ C. $\text{Mg (s)} / \text{Mg}^{2+}(\text{aq}) // \text{Zn}^{+2}(\text{aq}) / \text{zn (S)}$
 ii. $\text{Mg (S)} / \text{Mg}^{+2}(\text{aq}) // \text{Zn (S)} / \text{Zn}^{+2}(\text{aq})$ D. $\text{Zn}^{2+}(\text{aq}) / (\text{Zn(S)} // \text{Mg}^{2+}(\text{aq}) / \text{Mg (s)})$

353. A voltaic cell is constructed with pb (s) & Pb (NO₃)₂ (aq) for one half cell and Zn is a better reducing agent then pb and the salt bridge contains an aqueous solution of KNO₃ which of the following is not true.

- A. Zn is a positive electrode
 B. Pb is a cathode
 C. The NO₃⁻ ion moves to the Zn /Zn⁺² compartments
 D. The mass of Zn electrode decreases

Brief:

- ✓ From the given information the cell reaction is: $\text{Zn/Zn}^{+2} // \text{Pb}^{+2}/\text{Pb}$
- ✓ Zn is anode electrode (negative electrode)
- ✓ Pb is cathode electrode (positive electrode)
- ✓ NO₃⁻ (from salt bridge) migrate towards positive compartment (those gave electron, Zn /Zn⁺² compartments)
- ✓ K⁺ (from salt bridge) migrate towards negative compartment (those gained electron), $\text{Pb}^{+2} // \text{Pb}$ compartments)

354. When a current of 1.5A was passed for 30 minutes through a solution of trivalent metal, 1.071g of the metal was deposited at the cathode calculate the atomic mass of metal

- A. 32.13 B. 110 C. 115 D. 125

Brief:

355. Which of the following cannot be reduced by Fe? [E^0 for Fe⁺²/Fe = 0.45v)

- A. Ni⁺², E^0 for Ni²/Ni = -0.26V C. Mg⁺², E^0 for Mg⁺²/Mg = -2.37V
 B. Sn⁺⁴, E^0 for Sn⁺⁴/Sn⁺² = 0.15v D. Cl₂, E^0 for cl₂/cl⁻ = 1.36V

Brief:

- ✓ The spontaneity of the reaction is depends on the standard reduction potential of reacting species
- ✓ The more positive E^0 reduction potential, easily reduced (difficult to oxidize)
- ✓ The more negative E^0 reduction potential, easily oxidized (difficult to reduce)
- ✓ So, Fe can reduce other species those having more positive E^0 reduction potential.
- ✓ So, Ni⁺², Sn⁺⁴, Mg⁺² can be reduced by Fe while Cl₂ can't reduce by Fe.

Charge passed through the solution = $1.5 \times 30 \times 60$
 = 2700 coulombs.

$$\text{Mole of electric charge passed} = \frac{2700}{96500} \text{ faraday} = \frac{27}{965} F.$$

$$\therefore \text{equivalent of the metal deposited} = \frac{27}{965}.$$

Let the at. Wt. of the metal be x.

$$\therefore \text{eq. wt. of the metal} = \frac{x}{3}. \quad (\because \text{valency of the metal} = 3)$$

$$\therefore \text{weight of the metal deposited} = \text{no. of eq.} \times \text{eq. wt.}$$

$$= \frac{27}{965} \times \frac{x}{3} g. \therefore \text{as given, } \frac{27 \times x}{965 \times 3} = 1.071$$

$$x = 114.8 \text{ amu.}$$

356. Write down the equation that shows how the emf of the cell changes as the concentrations of chromium (III) and cerium (III) ions change for the reaction whose cell notation is

given below . $\text{Ce (s) / ce}^{+3} \text{ (aq) // cr}^{+3} \text{ (aq) / cr(s)}$

$$\text{A. } E_{\text{cell}} = E_{\text{cell}}^0 - \frac{0.0592}{3} \log \left[\frac{\text{ce}^{+3}}{\text{cr}^{+3}} \right]$$

$$\text{B. } E_{\text{cell}} = E_{\text{cell}}^0 + \frac{0.0592}{3} \log \left[\frac{\text{ce}^{+3}}{\text{cr}^{+3}} \right]$$

$$\text{C. } E_{\text{cell}} = E_{\text{cell}}^0 - 0.059 \log \left[\frac{\text{ce}^{+3}}{\text{cr}^{+3}} \right]$$

$$\text{D. } E_{\text{cell}} = E_{\text{cell}}^0 + \frac{0.0592}{3} \log \left[\frac{\text{cr}^{+3}}{\text{Ce}^{+3}} \right]$$

357. Two electrolysis baths are placed in series one both contains cupric sulphate and the second silver nitrate. If 1.6g of copper are produced by electrolysis. Calculate the mass of silver deposited? [At mass of Ag = 108 & ca= 64g /mol) .

A. 5.4g B. 4.8g C. 5.3g D. 3.2g

Brief:

- If equal amount of charge (Q) is passed through two different solutions, the amount of substance deposited/liberated (m) is proportional to their chemical equivalent weights (E).
- For copper cell,
- By Faradays first law, $m_{\text{Cu}} = Z_{\text{Cu}} \times Q_{\text{Cu}}$
 where,

m_{Cu} = Amount of copper deposited

Z_{Cu} = Electrochemical equivalent of copper

* Q_{Cu} = Charge passed through the copper cell

* For Silver cell, By Faradays first law,

* $m_{\text{Ag}} = Z_{\text{Ag}} \times Q_{\text{Ag}}$, where,

m_{Ag} = Amount of silver deposited

Z_{Ag} = Electrochemical equivalent of silver

Q_{Ag} = Charge passed through the silver cell

* Since, charge passed is same for both the cells.

○ $Q_{Cu} = Q_{Ag}$ So,

○ $m_{Cu}/m_{Ag} = Z_{Cu}/Z_{Ag}$

○ $Z_{Cu} = E_{Cu}/96500$

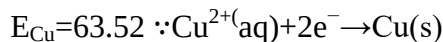
○ $Z_{Ag} = E_{Ag}/96500$

$m_{Cu}/m_{Ag} = Z_{Cu}/Z_{Ag}$

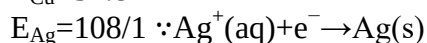
$m_{Cu}/m_{Ag} = E_{Cu}/E_{Ag} = Z_{Cu}/Z_{Ag}$

E_1 = chemical equivalent weight of copper E_2 = chemical equivalent weight of silver

E = Atomic Mass n – factor



$E_{Cu} = 31.8$



$m_{Cu}/m_{Ag} = E_{Cu}/E_{Ag}$

Hence, $1.6g/m_{Ag} = 31.8/108 \Rightarrow m_{Ag} = 1.6 \times 108/31.8$ $m_{Ag} = 5.4g$

358. The electrolyte in a lead storage battery is

A. H_2SO_4 B. KOH

C. HNO_3 D. $Ca(OH)_2$ E. NH_4Cl

Primary batteries					
type	Anode	Cathode	Electrolyte	Voltage	common applications
zinc-carbon (Leclanché)	zinc alloy	manganese dioxide	80% ammonium chloride and 20% zinc chloride	1.5	used in remote controls, flashlights, portable radios
zinc chloride	zinc	manganese dioxide	zinc chloride electrolyte	1.55	used in motorized toys, cassette and CD players, flashlights, portable radios
zinc-manganese dioxide	zinc	manganese dioxide	potassium hydroxide	1.5	best for use in motorized toys, cassette and CD players
zinc-silver oxide	zinc anode	silver oxide cathode	potassium hydroxide	1.6	used in hearing aids, watches, calculators
lithium-iron sulfide	lithium	iron sulfide	organic electrolyte	1.6	used in digital cameras, small appliances
lithium-manganese dioxide	lithium	manganese dioxide	organic electrolyte	2.8-3.2	used in digital cameras, small appliances
Secondary batteries					
lead-acid	lead	lead dioxide	sulfuric acid	12 (2v*6 cell) = 12v	used in automobiles, wheelchairs, children's electric vehicles, emergency power supplies
nickel-cadmium	cadmium	nickel dioxide	potassium hydroxide	1.3	used in power tools, cordless telephones, biomedical equipment
nickel-metal hydride	nickel alloy	nickel dioxide cathode	potassium hydroxide	1.2	used in smoke alarms, power tools, cellular telephones

359. How many electrons are entered to cathode of electrolytic cell by a current of 5A in 10 minutes?

A. 0.03

B. 50

C. 1.87×10^{22}

D. 6.02×10^{23}

Brief:

- * Given: Current $I=5\text{A}$
- * Time $t= 600\text{s}$
- * Charge on one electron, $Q = 1.6\times 10^{-19}\text{ C}$
- * Let the number of electrons be $= n$
- * We know that $I = Q/t$ and $Q = n\times e$
- * $\Rightarrow I = (n\times e)/t$
- * On substituting the values, $5\text{A} = \frac{n\times 1.6\times 10^{-19}}{600\text{s}}$
- * $\Rightarrow n = (5\text{A}\times 600\text{s})/1.6\times 10^{-19} = 1.87\times 10^{22}$
- * Therefore, the number of electrons flowing in 600 second when 5 A current flows through a wire is 1.87×10^{22}

360. How many moles of a trivalent metal would be deposited at a cathode during the electrolysis for the passage of 20A for 3 hour?

- A. 0.25 B. 0.50 C. 0.75 D. 2.25

Brief:

❖ Quantity of charge passed $= It =$

- * $20\text{A} \times 3 \times 60 \times 60 = 216000\text{ coulomb}$
- * On depositing one mole metal 96500C charge passing through the cell
- * $\text{Mass (m)} = \frac{MIt}{nF}$
- * $\frac{m}{M} = \frac{It}{nF}$
- * $n = \frac{It}{nF} = \frac{216000\text{ C}}{(3)(96500\text{C})} = 0.75$
- * $n =$

361. In which of the following cells a reversible electro chemical process can take place? (self-practice)

- A. Zinc carbon dry cell C. Lead storage battery
B. Alkaline battery D. copper (II) oxide cell

362. 0.2964 g of copper was deposited on passage of a current of 0.5 ampere for 30 minutes through a solution of copper sulphate. Calculate the atomic mass of copper. (1 faraday = 96500 coulomb)

- ❖ 63.56g/mol B. 67.22g/mol C. 98.12g/mol D. 18.54g/mol

❖ Quantity of charge passed

$$0.5 \times 30 \times 60 = 900\text{ coulomb}$$

$$900\text{ coulomb deposit copper} = 0.2964\text{ g}$$

96500 coulomb deposit copper = $0.2964 / (900 \times 96500) = 31.78 \text{ g}$

Thus, 31.78 is the equivalent mass of copper.

At. mass = Eq. mass $\times$ Valency = $31.78 \times 2 = 63.56$

363. Aqueous solution of CuSO_4 is electrolytes using Cu- electrodes which of the following statement are true. (*Refer on your text book, page 79.*)

- A. The PH of the solution increases
- B. The anode increases in mass and the cathode decreases in mass
- C. The blue colour of the solution turns to colorless
- D. O_2 is Liberated at anode and H_2 is Librated at cathode

364. During the electrolysis of Brine the substance produce at the Anode is

- A. O_2
- B. Na
- C. Cl_2
- D. H_2

Brief:

- ❖ During electrolysis of concentrated NaCl solution (called brine)
- ❖ There are two anions (Cl^- and OH^-) and cations (Na^+ and H^+) in the cell, but according to preferential discharge concept, Cl^- and H^+ are easily discharge at anode and cathode electrode respectively.

365. Which one of the following equations shows the relationship between standard Gibbs free energy and equilibrium constant ?

- A. $K = \Delta G^0$
- B. $\Delta G^0 = -RT \ln K$
- C. $K = RT \ln \Delta G^0$
- D. $\Delta G^0 = RT \ln K$

Brief:

- The standard change in free energy, ΔG^0 , for a reaction is related to its equilibrium constant, K, by the equation
- $\Delta G^0 = -RT \ln K$. When $\Delta G^0 < 0$, $K > 1$, and the reaction is product-favored at equilibrium.
- When $\Delta G^0 > 0$, $K < 1$, and the reaction is reactant-favored at equilibrium.

366. Which serves as anode in Nickcl – cadmium battery? (refe question number 242)

- A. Cd(S)
- B. $\text{NiO}_2(\text{S})$
- C. $\text{PbO}_2(\text{S})$
- D. $\text{MnO}_2(\text{S})$
- E. Ni

367. Which one of the following does not affect preferential discharges ?

- A. Nature of the electrode
- C. Position of ions in electrochemical series
- B. Acidity of ions
- D. Concentration of the ions in the electrolyte

Brief:

- ✓ The preferential discharge theory states that if more than one type of ion is attracted towards a particular electrode, then the one discharged is the ion that requires the least energy.
- ✓ There are three criteria (factors) to select species those are easily discharge

368. The oxidation state of “Ni” ion in the complex $K_4(NiCN_4)$ is (self-practice)

- A. +4 B. +3 C. +2 D. 0 E. +1

369. What products are formed during the electrolysis of aqueous solution of the salt

$AgNO_3$

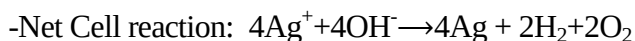
- A. Ag, O_2 , NO_3 B. Ag, O_2 , HNO_3 C. Ag, O_2 , H_2 D. H_2 , O_2 , NO_3

Brief:

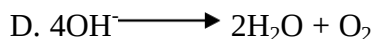
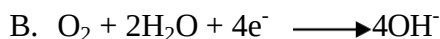
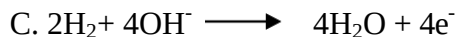
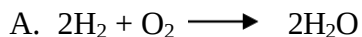
- * When the electrode of the same material is used for example Ag. The metal itself gets oxidized and reduced on an electrode.



When platinum electrodes are used, Ag^+ from solution is reduced and deposited at cathode whereas O_2 is produced at the anode.



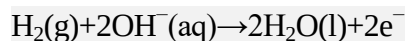
370. In hydrogen – oxygen fuel cell the cathode reaction is



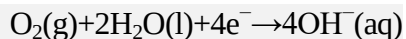
Brief:

- * A fuel cell is essentially a battery, but it differs in operating with a continuous supply of energetic reactants, or fuel.
- * In its simplest form, a hydrogen-oxygen fuel cell consists of an electrolyte solution, such as potassium hydroxide solution, and two inert electrodes.
- * Hydrogen and oxygen gases are bubbled through the anode and cathode compartments
- * In general, the chemical energy (heat of combustion) of fossil fuels (coal, gas, or oil) is first used for converting water into high pressure steam and then used to run a turbine to produce electricity.
- * In fuel cell, the chemical energy of combustion of fuels like H_2 , CH_4 , CH_3OH etc. are directly converted into electrical energy.
- * In hydrogen-oxygen fuel cell, hydrogen and oxygen are bubbled through porous carbon electrodes into concentrated aqueous sodium hydroxide solution. Here, finely divided **platinum or palladium** metal is used as catalyst.

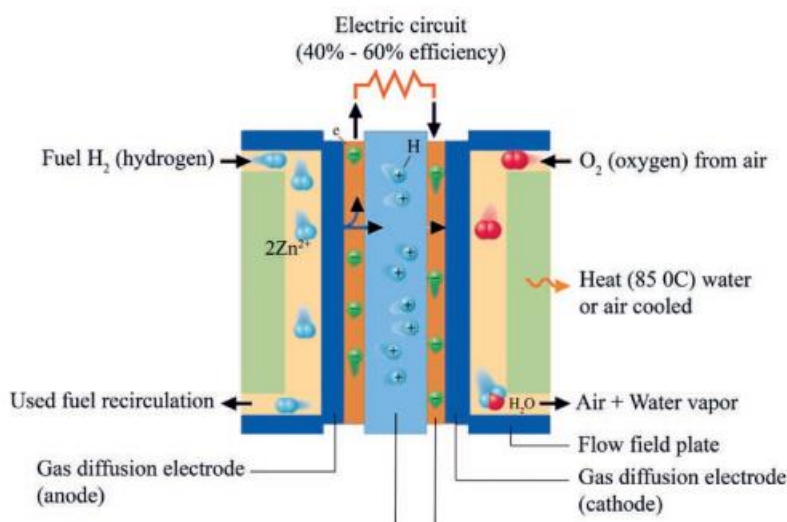
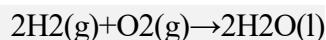
-At the anode:



-At the cathode:



-Overall reaction is:



Hydrogen – oxygen fuel cell.

Working Principle of Hydrogen – oxygen fuel cell

- ✓ At the anode, hydrogen (H_2) is split into protons and electrons.
- ✓ The electrons travel through an external circuit, generating electricity.
- ✓ The protons move through the electrolyte to the cathode, where they react with oxygen (O_2) to form water (H_2O) as a byproduct.

Advantages

- ✓ High efficiency
- ✓ Low environmental impact (emits only water in hydrogen fuel cells)
- ✓ Silent operation

Applications

- Electric vehicles (EVs)
- Backup power systems
- Spacecraft and submarines

371. For the oxidation of ammonia $4\text{NH}_3 + 3\text{O}_2 \longrightarrow 2\text{N}_2 + 6\text{H}_2\text{O}$, $\Delta G = 1356\text{KJ}$

The reaction can be run in a galvanic cell calculate the standard cell potential

- A. 2.17V B. 1.17V C. 3.51V D. 7.11V E. None

Brief:

The cell potential drops as the reactants are consumed and the reaction approaches equilibrium

$$\Delta G = -n\mathcal{F}E_{\text{cell}} \Rightarrow \text{when } \Delta G = 0, E_{\text{cell}} = 0$$

$$\Delta G = \Delta G^\circ + RT \ln Q$$

$$E = E^\circ - \frac{0.0257 \text{ V}}{n} \ln Q \quad \text{or} \quad E = E^\circ - \frac{0.0592 \text{ V}}{n} \log Q$$

Then, calculate $E_{\text{cell}}^\circ, E_{\text{cell}}$.

372. In lead – storage battery the product formed at the anode & cathode respectively

- A. $\text{PbO}_2, \text{PbSO}_4$ B. $\text{PbSO}_4, \text{PbSO}_4$ C. $\text{PbSO}_4, \text{PbO}_2$ D. $\text{PbSO}_4, \text{H}_2\text{SO}_4$

Brief:

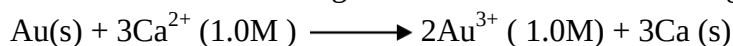
* The discharging and charging reactions are:

Oxidation: $\text{Pb} + \text{SO}_4^{2-} \rightarrow \text{PbSO}_4 + 2\text{e}^-$ (charging)

Reduction: $\text{PbO}_2 + 4\text{H}^+ + \text{SO}_4^{2-} + 2\text{e}^- \rightarrow \text{PbSO}_4 + 2\text{H}_2\text{O}$ (discharging)

NB: PbSO_4 is the common product in both oxidation and reduction reaction.

373. Which of the following is not true about the reaction given below at 25°C



$$E^\circ_{\text{Ca}^{2+}/\text{Ca}} = -2.87\text{V}, E^\circ_{\text{Au}^{3+}/\text{Au}} = +1.50$$

- A. The number of electrons involved are six
 B. The standard free energy (ΔG°) = $2.53 \times 10^3 \text{ KJ/mol}$
 C. The reaction favours the reactants at equilibrium
 D. The reaction is spontaneous at 25°C

Brief:

✓ Given The reaction: $2\text{Au(s)} + 3\text{Ca}^{2+} (1.0\text{M}) \longrightarrow 2\text{Au}^{3+} (1.0\text{M}) + 3\text{Ca(s)}$

✓ $\Delta G^\circ = -RT \ln K$

✓ $E^\circ_{\text{cell}} = E^\circ(\text{reduction}) + E^\circ(\text{oxidation})$

$$\Delta G = E_{\text{cell}} - \frac{0.0592}{6} \log \frac{[\text{Au}^{3+}]}{[\text{Ca}^{2+}]}$$

374. Without doing any of calculation, what would be the standard half-cell reduction potential of

lead metal if the measured E_o of the cell: $\text{Pb(s)} | \text{Pb}^{2+}(1\text{M}) || \text{H}^+(1\text{M}) | \text{H}_2(1\text{atm}) | \text{Pt(s)}$ is + 0.13

- A. 0.03v B. -0.13v C. 0.13v D. 0.00v

Brief:

✚ The anode where oxidation occurs is listed first. The double bar refers to the salt bridge.

- ✚ The right side of the cell diagram refers to the cathode where reduction occurs.
- ✚ Since one of the two half-cell reaction is 0.00v, the standard reduction potential value of the other cell reaction is the reverse of the cell potential.

- ✓ Anode: $\text{Pb(s)} \rightarrow \text{Pb}^{2+}(\text{aq}) + 2\text{e}^-$ $E^\circ_{\text{anode}} = ?$
- ✓ Cathode: $2\text{e}^- + 2\text{H}^+ \rightarrow \text{H}_2(\text{g})$ $E^\circ_{\text{cathode}} = 0.00 \text{ V}$ (by definition)
- ✓ $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 0.00 \text{ V} - E^\circ_{\text{anode}} = +0.13 \text{ V}$
- ✓ $E^\circ_{\text{anode}} = 0.00 \text{ V} - 0.13 \text{ V} = -0.13 \text{ V}$

375. How can gold be plated on to spoon made of iron metal
- Electricity to iron anode in a solution of gold ions
 - Electricity to gold anode in a solution of iron ions
 - A solution of iron ions to react with a piece of gold
 - Electricity to iron cathode in a solution of gold ions

Brief:

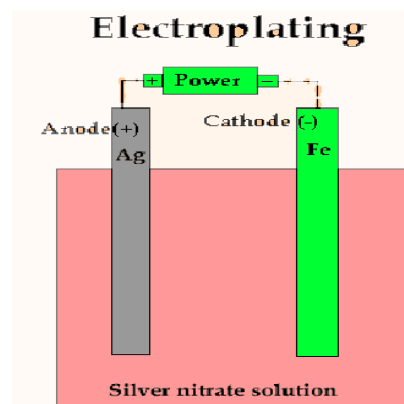
- ❖ Electroplating involves the deposition of a layer of metal on top of another metal.
- ❖ Often the silverware you see on dinner tables is cheap iron coated with a thin layer of silver. This is done by a process called electroplating in an electrolytic cell. A normal electrolytic cell is used for electroplating with the following conditions.

- ✓ The metal to be coated is placed at the negative terminal (cathode)
- ✓ The metal we wish to coat with is placed at the positive terminal (anode)
- ✓ The electrolyte is a solution of positive metal ions of the metal we wish to coat with.
- ✓ The concentration of the metal ion in the electrolyte stays constant during the electrolysis process

✓ As is shown on the right, If we wish to coat iron metal with a layer of silver then the electrolyte used should contain silver ions and the anode should be silver metal, while the iron is connected to the negative terminal.

✓ Electrons are pumped out of the silver metal in effect corroding it and releasing silver ions into the electrolyte solution. $\text{Ag(s)} \Rightarrow \text{Ag}^+(\text{aq}) + \text{e}^-$

✓ The positive ions migrate over to the negative electrode (the metal we wish to coat) and regain their electrons to form silver metal once again. $\text{Ag}^+(\text{aq}) + \text{e}^- \Rightarrow \text{Ag(s)}$

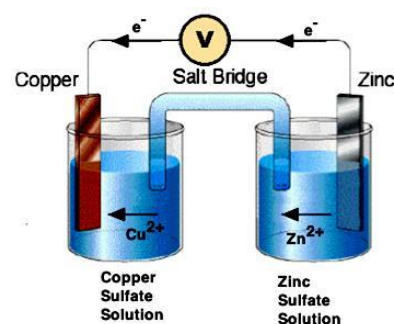


376. During manufacturing of sodium metal in Down cell process, CaCl_2 is added to molten NaCl to decrease the melting point of NaCl , because .
- CaCl_2 has lower melting point than NaCl
 - Impurity decrease the melting point
 - NaCl has lower melting point
 - Addition of CaCl_2 produce H_2O

Brief:

- ❖ Because with the presence of calcium ions, the sodium ions are farther from each other and moving more randomly in respect to one another in the molten solution. Thus the sodium ions need to be cooler (lower temperature, lower energy) to solidify (freeze) or, conversely, melt.

377. The diagram shows an electric current being produced by a galvanic cell. Which statement is NOT correct? (self-practice)



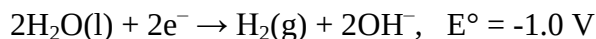
- Zinc ions are produced at the zinc electrode.
- The electrons flow through the wires from the zinc to the copper electrode.
- The salt bridge maintains a balance between the ions in both parts of the cell.
- Copper ions are produced at the copper electrode.

378. During the electrolysis of a concentration aqueous solution of NaCl , what substance is formed at the cathode?

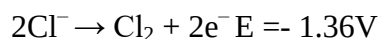
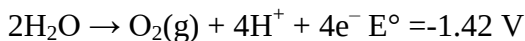
- Hydrogen
- Oxygen
- Chlorine
- Sodium

Brief:

- ❖ Sodium chloride is dissociated and exists as sodium and chloride ions in aqueous solution. Electrolysis of sodium chloride is easier in aqueous solution. But the water itself can undergo reduction and oxidation reactions at different potentials. So, the substance that is oxidized or reduced is not sodium and chloride ions alone, but it may involve the water molecule also.
- ❖ Two competing reactions are possible at both the cathode and anode.
- ❖ Water can be reduced to hydrogen gas or sodium ions reduced to sodium metal.

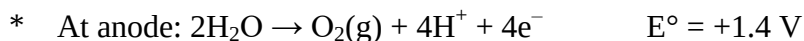
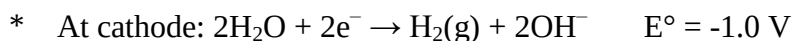


- ❖ At anode: oxidation reaction: at $\text{pH} = 7$. Water can be oxidized to oxygen or chloride ion oxidized to chlorine molecules.



A. *Very Dilute Aqueous Sodium Chloride Solution*

- ❖ Water has very low conductivity, and the small amount of ionic sodium chloride helps the ionic conductivity of the solution. In small concentrations, the electrolysis of water becomes more predominant yielding hydrogen at the cathode and oxygen at the anode.

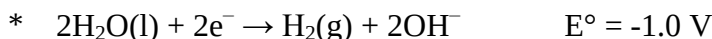


- ❖ The net reaction of electrolysis of very dilute aqueous sodium chloride is given as,



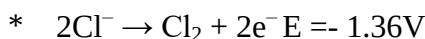
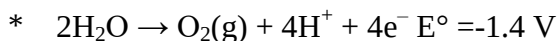
B. **High Concentration of Sodium Chloride**

- ❖ At cathode: reduction reaction: at pH =7

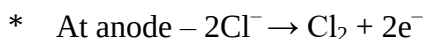
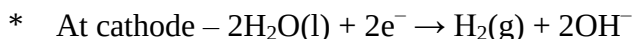


- ❖ The reduction (over) potential of water being more positive than the reduction of sodium ions, water is decomposed to liberate hydrogen at the cathode.

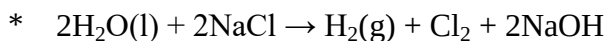
- * At anode: oxidation reaction: at pH =7



- ❖ The net reaction taking place is given as follows:



✚ The overall reaction is as follows:



✚ At high salt concentration, hydrogen and chlorine is the product of electrolysis with sodium hydroxide as a by-product.

- ❖ Hence, the product of electrolysis of aqueous sodium chloride can be anything between,

i) Sodium metal or hydrogen gas at the cathode

ii) Chlorine or oxygen gas at the anode

- ❖ With a side product of sodium hydroxide resulting from the reaction of sodium and water, the product of electrolysis depends on the concentration of sodium chloride aqueous solution.

379. What is the difference between electrolytic cell and galvanic cell?
- A. In galvanic cell non – spontaneous chemical reactions are takes place
 - B. In electrolytic cell non – spontaneous chemical reactions are takes place
 - C. Anode is positive whereas cathode is negative in galvanic cell
 - D. Electrolytic cells convert chemical reaction into electrical energy

Brief:

Galvanic Cell	Electrolytic cell
A Galvanic cell converts chemical energy into electrical energy.	An electrolytic cell converts electrical energy into chemical energy.
It is spontaneous and is responsible for the production of electrical energy.	The redox reaction is not spontaneous and electrical energy has to be supplied to initiate the reaction.
The two half-cells are set up in different containers, being connected through the salt bridge or porous partition.	Both the electrodes are placed in a same container in the solution of molten electrolyte.
Here the anode is negative and cathode is the positive electrode. The reaction at the anode is oxidation and that at the cathode is reduction.	Here, the anode is positive and cathode is the negative electrode. The reaction at the anode is oxidation and that at the cathode is reduction.
The electrons are supplied by the species getting oxidized. They move from anode to the cathode in the external circuit.	The external battery supplies the electrons. They enter through the cathode and come out through the anode.

$$\text{Zn (s)} \rightarrow \text{Zn}^{2+} (\text{aq}) + 2\text{e}^- \quad E^\circ = -0.76$$

$$\text{Cu}^{2+} (\text{aq}) + 2\text{e}^- \rightarrow \text{Cu (s)} \quad E^\circ = -0.34$$

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 0.34 \text{ V} - (-0.76 \text{ V}) = 1.10 \text{ V}$$

*Given that $E^\circ_{\text{galvanic}} > 0$ and $\Delta G^\circ = -nFE$, ΔG° for a galvanic cell must be negative and **spontaneous**!*

$$2\text{Cl}^- \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^- \quad E^\circ = 1.36\text{V}$$

$$2\text{Na}^+ + 2\text{e}^- \rightarrow 2\text{Na (l)} \quad E^\circ = -2.71\text{V}$$

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = -2.71\text{V} - (1.36\text{V}) = -4.07\text{V}$$

*Given that $E^\circ_{\text{electrolytic}} < 0$ and $\Delta G^\circ = -nFE$, ΔG° for an electrolytic cell must be positive and **nonspontaneous**!*

380. In a Voltaic cell, electrons from the _____ to the _____.

- A. Salt bridge, anode
 B. anode, Salt bridge
 C. Salt bridge, cathode
 D. anode, cathode

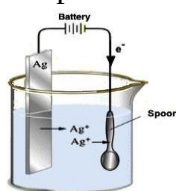
Brief:

Anode	Cathode	Electron flow	Site of oxidation	Site of reduction
-Ve	+Ve	Anode to cathode	Anode	Cathode

QUESTION NUMBER 383-

381. When electroplating nickel spoon with silver, which one of the following equations represents what happens at the anode? (self-prac

- A. $\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^-$
 B. $\text{Ag} \rightarrow \text{Ag}^+ + \text{e}^-$
 C. $\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$
 D. $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$



382. What does Faraday's First Law of Electrolysis state?

- A) The amount of substance deposited or liberated is directly proportional to the current and time
 B) The current is directly proportional to the voltage.
 C) The total charge is conserved in an electrolytic process.
 D) The resistance of the electrolyte affects the amount of substance deposited.

383. What is a common application for sacrificial anodes?

- A. In batteries to increase voltage.
 B. In pipelines and tanks to prevent corrosion.
 C. In electrical circuits to improve conductivity.
 D. In photovoltaic cells to enhance efficiency

384. In electrolytic conduction, what is the effect of increasing the temperature on the conductivity of an electrolyte solution?

- A. Conductivity decreases.
 B. Conductivity remains unchanged.
 C. Conductivity increases.
 D. Conductivity fluctuates.

385. What happens to the $\text{Cr}^{3+}(\text{aq})$ ion when it is converted to $\text{CrO}_4^{2-}(\text{aq})$ ion?

- A. Its oxidation number decreases and it undergoes reduction.
 B. Its oxidation number decreases and it undergoes oxidation.
 C. Its oxidation number increases and it undergoes reduction.
 D. Its oxidation number decreases and it undergoes reduction.

386. Which of the following statements is correct?

- A. The oxidizing agent is always oxidized in a redox reaction
- B. A reducing agent loses electrons in a redox reaction
- C. The reducing agent is reduced in a redox reaction
- D. The oxidizing agent always loses electrons

387. Which of the following statements is TRUE regarding sacrificial anodes?

- A. They are used only in freshwater applications.
- B. They can be made from metals such as magnesium, zinc, or aluminum.
- C. They require an external power source to operate.
- D. They function through the application of an electric current.

388. What is the primary function of a sacrificial anode in corrosion protection?

- A. To increase the rate of corrosion in metal structures.
- B. To act as a protective barrier against environmental factors.
- C. To sacrifice itself to protect another metal from corrosion.
- D. To enhance the electrical conductivity of the protected metal.

389. Which of the following factors does not affect the electrode potential of an electrode?

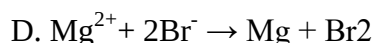
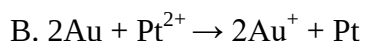
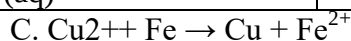
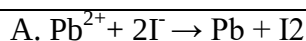
- A. Nature of the electrode (metal)
- B. Temperature of the solution
- C. Molarity of the solution
- D. Size of the electrode

390. What is the direction of flow of electrons in an electrolytic cell?

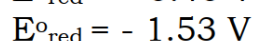
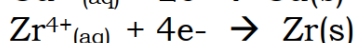
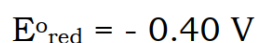
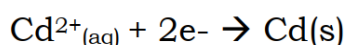
- A. Anode to cathode externally
- B. Anode to cathode internally
- C. Cathode to anode externally
- D. Cathode to anode in the solution

391. Which one of the following reactions is spontaneous (in the direction given) under standard electrochemical conditions? Given the following information;

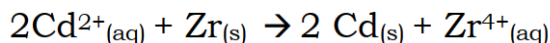
$\text{Au}^+ + \text{e}^- \rightarrow \text{Au(s)}$	$E^\circ_{\text{red}} = 1.83 \text{ V}$	$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb(s)}$	$E^\circ_{\text{red}} = -0.13 \text{ V}$
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$	$E^\circ_{\text{red}} = 0.34 \text{ V}$	$\text{Pt}^{2+} + 2\text{e}^- \rightarrow \text{Pt(s)}$	$E^\circ_{\text{red}} = 1.20 \text{ V}$
$\text{I}_2(\text{s}) + 2\text{e}^- \rightarrow 2\text{I}^-(\text{aq})$	$E^\circ_{\text{red}} = 0.54 \text{ V}$	$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe(s)}$	$E^\circ_{\text{red}} = -0.41 \text{ V}$
$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Mg(s)}$	$E^\circ_{\text{red}} = -2.38 \text{ V}$	$\text{Br}_2(\text{l}) + 2\text{e}^- \rightarrow 2\text{Br}^-(\text{aq})$	$E^\circ_{\text{red}} = 1.07 \text{ V}$



392. Given the following two half reactions,



393. Determine E°_{cell} and the spontaneity of the following reaction;



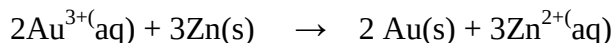
A. +1.13 V and not spontaneous

C. +1.13 V and spontaneous

B. -1.13 V and not spontaneous

D. -1.13 V and spontaneous

394. Consider an electrochemical cell where the following reaction takes place:



Which of the following is the correct cell notation for this cell?

A. $\text{Au} | \text{Au}^{3+} || \text{Zn}^{2+} | \text{Zn}$

(C) $\text{Au}^{3+} | \text{Au} || \text{Zn} | \text{Zn}^{2+}$

B. $\text{Zn} | \text{Zn}^{2+} || \text{Au}^{3+} | \text{Au}$

(D) $\text{Zn} | \text{Au}^{3+} || \text{Au} | \text{Zn}^{2+}$

395. Corrosion of iron is retarded by _____.

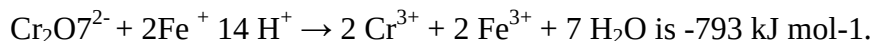
A. Both the presence of salts and low pH conditions

B. High pH conditions

C. Both the presence of salts and high pH conditions

D. The presence of salts

396. The standard reaction Gibbs energy for the electrochemical reaction



Calculate the standard cell potential. $F = 96500\text{ C}$

A. +1.37 V

B. +4.11 V

C. +2.74 V

D. +2.05 V

397. Why is direct current (DC) used in electroplating?

A. To alternate the direction of ion flow

C. To speed up the process

B. To create a uniform metal coating

D. To prevent chemical reactions

398. Which is the active material present on the negative plate in a lead-acid battery?

A. Lead dioxide

B. Lead peroxide

C. Spongy lead

D. Water

399. Which of the following statements regarding corrosion is true?

A. Corrosion does not depend on the reactivity of the metal

B. Presence of impurities does not affect the rate of corrosion

C. Strains in metals affect the rate of corrosion

D. Presence of electrolytes does not affect the rate of corrosion

400. Which of the following statements is false?

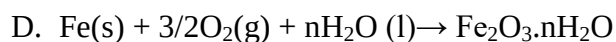
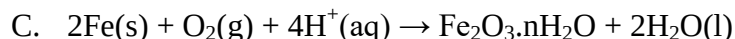
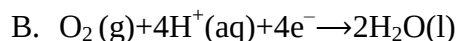
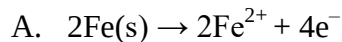
A. Salt water decelerates the rate of corrosion

B. Magnesium is more active than iron

C. During galvanization, $\text{ZnCO}_3 \cdot \text{Zn}(\text{OH})_2$ is formed which prevent further corrosion

D. Anti-rust solutions are used in car radiators to prevent rusting of iron parts of the engine

401. Which of the following is the overall reaction of rusting?



402. What is the purpose of the salt bridge in an electrochemical cell?
- To increase the cell potential
 - To allow the movement of electrons between half-cells
 - To maintain electrical neutrality by allowing ion flow
 - To prevent the mixing of solutions
404. A spoon is made the cathode in an electroplating apparatus containing AgNO_3 solution. How many grams of silver will be plated on the spoon if a current of **10A** is passed through the apparatus for **965** seconds?
- 5.4 g
 - 10.8 g
 - 2.7 g
 - 1.35 g
405. An electrochemical cell constructed for the reaction:
- $$\text{Ni}^{2+}(\text{aq}) + \text{M}(\text{s}) \rightarrow \text{Ni}(\text{s}) + \text{M}^{2+}(\text{aq})$$
- has $E_{\text{cell}}^0 = 0.16\text{V}$. The standard reduction, E_{red}^0 potential for Ni^{2+}/Ni is -0.25V . What is the standard reduction potential for M^{2+}/M ?
- -0.09 V
 - $+0.410\text{ V}$
 - -0.410 V
 - $+0.09\text{ V}$
406. In the electrolysis of dilute aqueous solution of Copper sulphate, CuSO_4 using inert electrodes, which is incorrect?
- Anode Reaction: $4\text{OH}^-(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g}) + 4\text{e}^-$
 - Cathode Reaction: $4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2(\text{l})$
 - The pH of the solution decreases
 - The concentration of copper (II) sulfate solution is decreased
407. Which of the following is **incorrect** about a galvanic cell represented by the cell notation:
- $$\text{Zn}(\text{s}) / \text{Zn}^{2+}(1.0\text{ M}) // \text{Ag}^+(1.0\text{ M}) / \text{Ag}(\text{s})$$
- Electrons will flow through the external circuit from the zinc to silver electrode
 - Reduction occurs at the zinc electrode as the cell operates
 - The mass of the zinc electrode will decrease as the cell operates
 - The concentration of Ag^+ will decrease as the cell operates
408. One of the basic differences between a fuel cell and a battery is:
- a fuel cell is rechargeable, while a battery is not.
 - a fuel cell does not store chemical energy, but a battery can store chemical energy.
 - both a fuel cell and a battery are rechargeable.
 - Both are 2ry cell

409. Which of the following statements is TRUE regarding sacrificial anodes?
- A. They are used only in freshwater applications.
 - B. They can be made from metals such as magnesium, zinc, or aluminum.
 - C. They require an external power source to operate.
 - D. They function through the application of an electric current.
410. What is the primary function of a sacrificial anode in corrosion protection?
- A. To increase the rate of corrosion in metal structures.
 - B. To act as a protective barrier against environmental factors.
 - C. To sacrifice itself to protect another metal from corrosion.
 - D. To enhance the electrical conductivity of the p

UNIT NINE

9. INDUSTRIAL CHEMISTRY

Unit out come

➤ ***At the end of this unit, you will be able to***

- ❖ *List the renewable and non-renewable natural resources and appreciate their importance to the industry as raw materials.*
- ❖ *Identify which resource is for which chemicals industry in Ethiopia*
- ❖ *List some important chemicals used for the manufacture of valuable products.*
- ❖ *explain the major steps and raw materials used in the industrial production of ammonia, nitric acid, sulphuric acid, sodium hydroxide, sodium carbonate and aluminum sulphate and di-ammonium phosphate and other Nitrogen fertilizers; and*
- ❖ *List some important direct uses of NH_3 , HNO_3 , H_2SO_4 , and Na_2CO_3 and their uses as raw materials in the production of other products*
- ❖ *Discuss the important steps and chemicals used in the production of glass, ceramics, cement, sugar and paper and pulp*
- ❖ *Describe the important steps and chemicals used in the production of soaps and detergent, and beverage*

9.1. Introduction

9.1.1. Manufacturing Chemicals and Goods

The chemical industry involves the transformation of raw materials into chemical products used in various applications. It plays a crucial role in economic development by providing essential chemicals for agriculture, healthcare, and consumer goods.

Chemical	Raw Materials	Process
Nitric Acid (HNO ₃)	Ammonia (NH ₃), Oxygen, Water	Ostwald Process: Catalytic oxidation of ammonia to NO, oxidation of NO to NO ₂ , absorption of NO ₂ in water to produce HNO ₃
Sulfuric Acid (H ₂ SO ₄)	Sulfur, Oxygen, Water	Contact Process: Sulfur burning to form SO ₂ , catalytic oxidation of SO ₂ to SO ₃ , absorption of SO ₃ in water to form H ₂ SO ₄
Ammonia (NH ₃)	Nitrogen (N ₂), Hydrogen (H ₂)	Haber Process: N ₂ and H ₂ mixed in a 1:3 ratio, catalytic synthesis at high temperature and pressure, ammonia separation by liquefaction

Sugar Industry

- ✓ Raw Materials: Sugarcane, sugar beet
- ✓ Process: Extraction of juice → Purification using lime and carbonation → Evaporation to concentrate syrup → Crystallization and centrifugation → Drying and packaging
- ✓ By-Products: Molasses, bagasse, press mud

Leather Industry

- ✓ Raw Materials: Animal hides and skins
- ✓ Process: Preservation (salting, drying) → Soaking and liming → Dehairing and deliming → Tanning (chromium, vegetable, or synthetic tanning) → Dyeing, fatliquoring, and finishing
- ✓ Environmental Concerns: Wastewater treatment, heavy metal pollution

Paper and Pulp Industry

- ✓ Raw Materials: Wood, bamboo, recycled paper
- ✓ Process: Pulping (mechanical, chemical, semi-chemical) → Bleaching → Paper sheet formation → Drying and finishing
- ✓ Environmental Concerns: Deforestation, chemical waste disposal

Soap and Detergent Industry

- ✓ Raw Materials: Fats and oils (for soap), petrochemicals (for detergents), alkalis (NaOH, KOH)
- ✓ Soap: Saponification of fats and oils with alkali
- ✓ Detergent: Sulfonation and neutralization of petroleum-based surfactants
- ✓ Mixing with additives, drying, and packaging

9.2. Questions with Explanation

411. Which of the following is field of chemistry incorrectly matched with the field of study?
- A. Industrial chemistry.....About transformation of natural raw materials and their derivatives in to products that have beneficial to humanity
 - B. Analytical chemistry Deal about Obtaining, processing, and communication information about the composition and structure of matter.
 - C. Physical chemistry Studies about how matter behaves on a molecular and atomic level.
 - D. Organic chemistry Studies about carbon-containing compounds like carbonates

Brief:

1. Organic chemistry: Studies carbon-containing compounds.
2. Inorganic chemistry: Focuses on compounds such as metal and mineral.
3. Physical chemistry: Incorporates physics into chemistry.
4. Biochemistry: Studies the chemical processes within living things.
5. Analytical chemistry: Focuses on the structure and composition of matter.
6. Industrial chemistry: Transformation of natural raw materials in to important products.

412. One of the following is the characteristics of chemical industry
- E. They are use naturally available raw material for more atom economy.
 - F. There is chemical reactions that transform raw materials in to products
 - G. They consume relatively small quantities of energy during the manufacturing
 - H. There should be safe operation methods in their manufacturing processes.

Explanation: Chemical industries are commonly characterized as

- ✓ Natural/artificial raw materials
- ✓ Coast consumption
- ✓ Material transformation
- ✓ Safe operation
- ✓ Quality control test.
- ✓ Unless controlled, they release to the environment as pollutant.

413. The group consists of non-renewable organic resources
- | | |
|------------------------------|-------------------------------|
| A. Water, air and minerals | C. wood, water and natural |
| B. natural gas, oil and coal | D. pasturessand, air and clay |

Brief:

- ❖ Renewable resources cannot be depleted over time.
Renewable resources include sunlight, water, wind and also geothermal sources such as hot springs and fumaroles. common sources of renewable energy:
- | | |
|----------------------|---------------|
| B. Solar energy. | E. Hydropower |
| C. Wind energy | |
| D. Geothermal energy | |

Non-renewable resources deplete over time.

Ocean energy, Bioenergy

414. Non-renewable will be gone forever once used, like coal, fuel, petroleum, oil, natural gases

Criteria	Non-renewable energy resources	Renewable energy resources
Repetitively used energy sources	One-time use	Can use again and again
Perpetual sources	Vanish one day	Available for perpetual use
Eco-friendly sources	Emits gasses and pollute the environment	Do not emit gases and does not pollute the environment
Availability	Limited quantity	Unlimited quantity.
Production cost	High	Low
Maintenance cost	Low	High
Production quantity	Large	Less

415. Which of the following component of environment is composed of different gases namely N_2 , O_2 , CO_2 , Ne, Ar, Kr and Xe?

A. Hydrosphere

C. Biosphere

Brief:

B. Lithosphere

D. Thermosphere

Atmosphere: The earth's atmosphere composed of different gases namely N_2 , O_2 , CO_2 , Ne, Ar, Kr

Hydrosphere: Ocean water which amounts to about 1.5×10^{21} liters contains about 3.5 percent by mass dissolved material.

Lithosphere: The earth's crust in the form of mineral ores, carbon and hydrocarbons, coal, natural gas and crude petroleum.

Biosphere: Vegetation and animals contribute raw materials (Oils, fats, waxes, resins, sugar, natural fibers and leather.

416. The solubility ammonia in water at 25° is 34% (w/w) implies that

A. 66g of ammonia is dissolving in 100g of water.

B. 34g of ammonia is dissolving in 100g of solution.

C. 34g of water is dissolving in 66g of ammonia.

D. 34g of ammonia is dissolving in 66g of water.

417. Which of the following are the two most important methods of ammonia manufacturing given the reaction respectively?



A. Heating of ammonium salt, Habber process

B. Habber process, Heating of ammonium salt

C. Solvey process, Habber process

D. Habber process, Down cell process

418. Which of the following is true regarding nitric acid?

A. It is a strong reducing agent

C. It is a strong oxidizing agent

B. Its basicity is 2

D. It is non-planar in gaseous state

Brief:

- ❖ It is a strong oxidizing agent. It has a polar O–H bond. This bond breaks to donate the H^+ ion which is why it is a strong oxidizing agent and a strong acid.
- ❖ Since there is only one cleavable O–H bond, the basicity of nitric acid is unity (one HNO_3 molecule can donate only one H^+ ion).
- ❖ It exists as a planar molecule in gaseous state
- ❖ Metals like Cr, Al, Pt, Au, etc do not dissolve in nitric acid because of the formation of passive film of oxide on its surface.

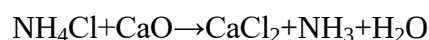
419. What is the role that calcium oxide in the preparation of ammonia gas?

- A. It is used as a reactant in the production of intermediates.
- B. It is used as a catalyst in the production of ammonia
- C. It is used as a drying agent in the production of ammonia
- D. It is used as a reactant and drying agent in the production of ammonia

Brief:

It is used as a reactant and drying agent in the production of ammonia.

The reaction to produce ammonia using ammonium chloride and calcium oxide is given below.



Water is also produced in the reaction and is present in the form of moisture which needs to be separated from the ammonia gas. Calcium oxide or quicklime not only helps in the production of ammonia gas but also acts as a drying agent by absorbing all the moisture that is released in this reaction. Hence calcium oxide is both a reactant and a drying agent in the experimental manufacturing of ammonia gas.

420. Which of the following is NOT correct about the downward displacement of air method used for drying and collecting ammonia gas?

- A. Ammonia gas is lighter than air and hence collected by the downward displacement of air
- B. Ammonia is not collected over water since it is highly soluble in water
- C. The drying agent used is calcium oxide.
- D. Oxygen is somewhat lighter than air; it is gathered by the downward displacement by air.

Brief:

- ❖ Since ammonia gas is lighter than air, it is collected by downward displacement of air.
- ❖ Ammonia gas is collected in an inverted gas jar by the downward displacement of air because ammonia is lighter than air (V.D. of ammonia is 8.5 and that of air is 14.4).

- ❖ It is highly soluble in water as it cannot be collected over water.

421. Why is oxygen collected by downward displacement of air?

- A. It can be collected under with fear of excessive dilution.
- B. It can be collected by downward displacement of air since oxygen is slightly heavier than air
- C. Water displacement is a method of collecting gases that are soluble in water.
- D. Oxygen is somewhat heavier than air, it is gathered by the downward displacement of water rather than by air.

Brief:

- ✓ Oxygen is collected by the downward displacement of water and not over air.
- ✓ Downward displacement of water is used to collect oxygen.
- ✓ Any gas be collected by this method and the gas pushes the water down as it is lighter and do not react with water
- ✓ Oxygen is somewhat heavier than air, it is gathered by the downward displacement of water rather than by air.

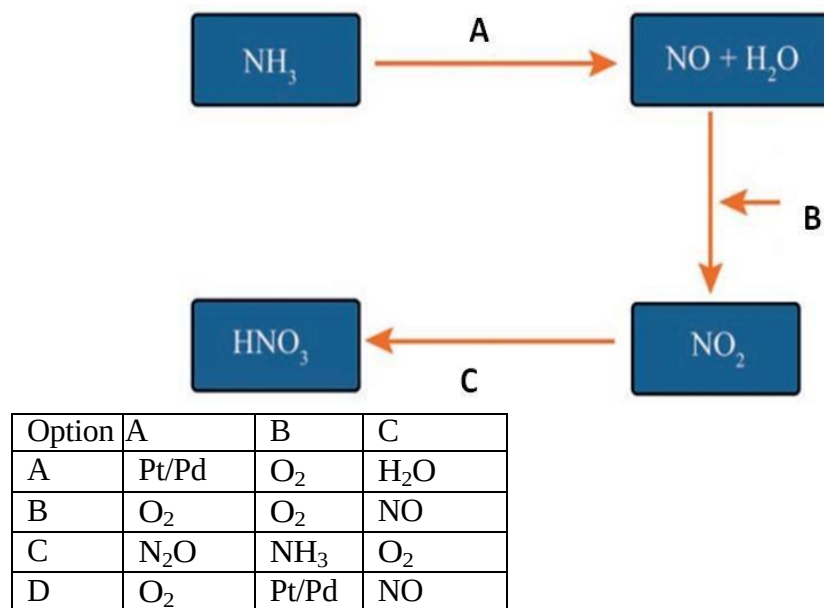
422. Which of the following is NOT suitable for reaction condition in manufacturing process of ammonia:

- A. The phenolphthalein present will indicate a basic solution (pink in color)
- B. Apply optimum Temperature for more yield of ammonia
- C. Increasing of Pressure forwarding the reaction to the product side
- D. Decreasing the pressure forwarding the reaction to the product side

Brief:

- ❖ The optimum condition are used to provides high amount of ammonia
- ❖ Ammonia is favored by high pressure (lower volume) and low (not too!) temperature
- ❖ Due to the exothermic nature of the reaction, the steady increase in heat over a period of time causes the gradual decomposition of ammonia. To prevent this from happening, the iron in the system is coupled with K_2O , CaO , and Al_2O_3 in order to catalyze the reaction and increase its speed. The iron adsorbs the nitrogen and hydrogen onto its surface, hence, lowering the amount of critical energy required to cleave the covalent bonds of nitrogen.

423. High temperature conditions are required in order to break the strong triple bond of the nitrogen molecules. But, due to the exothermic nature of the reaction, the steady increase in heat over a period of time causes the gradual decomposition of ammonia. Given the steps in industrial manufacturing process of nitric acid (Ostwald process), which of the following are correctly matched with their chemicals/catalysts used in each steps.



424. Which of the following is Nitrogenous fertilizer

- A. Copper phosphate
B. Potassium sulphate

- C. Potassium chloride
D. Urea

Brief:

➤ Nitrogenous fertilizers are containing nitrogen as major components

- ❖ Nitrate fertilizers
- ❖ Ammonium fertilizers
- ❖ Nitrate and ammonium fertilizers
- ❖ Amide fertilizers

Tip: (i) Urea-

It is slightly acidic in nature. Urea application to soil creates a small loss of calcium from the soil.

Urea is highly concentrated nitrogenous fertilizer containing 46% nitrogen.

(ii) Calcium Cyanamid- It is non-leachable, synthetic organic fertilizer. Its basicity (alkalinity) is three times as high as that of sodium nitrate and advisable for use on acid soils.

425. Find the correct match of the columns in element supplied by fertilizer of fertilizers.

- | | |
|------------------|--------------------|
| (i). Nitrogen | A. Super phosphate |
| (ii). Phosphorus | B. Potash |
| (iii). Potassium | C. Urea |

- A. (i) - A, (ii) - C, (iii) - B
B. (i) - A, (ii) - B, (iii) - C
C. (i) - B, (ii) - C, (iii) - A
D. (i) - C, (ii) - A, (iii) - B

426. Which one of the following nutrients is not available in fertilizers?

- A. Nitrogen B. Phosphorus C. Iron D. Potassium

Brief:- Fertilizers - Applied to soils or the plant tissues for their growth.

- ✓ Nitrogen - For leaf growth
- ✓ Potassium - For strong stem growth
- ✓ Phosphorous - For the development of roots, flowers, seeds and fruits.

427. Which of the following is/are NOT phosphatic fertilizers?

- A. Ammonium hydrogen phosphate
B. Mono hydrogen phosphate
C. Ammonium chloride
D. Di-ammonium phosphate

Brief:

-Ammonium dihydrogen phosphate (ADP), also known as monoammonium phosphate, is formed when a solution of phosphoric acid is added to ammonia until the solution is distinctly acidic.

-Diammonium phosphate is one of a series of water-soluble ammonium phosphate salts that can be produced when ammonia reacts with phosphoric acid.

-DAP is used as a fertilizer. When applied as plant food, it temporarily increases the soil pH, but over a long term the treated ground becomes more acidic than before upon nitrification of the ammonium.

-Hydrogen phosphate or monohydrogen phosphate is the inorganic fertilizer with the formula $[\text{HPO}_4]^{2-}$. Its formula can also be written as $[\text{PO}_3(\text{OH})]^{2-}$.

-Together with dihydrogen phosphate, hydrogen phosphate occurs widely in natural systems, Fertilizer, cooking.

427. Which of the following is true about the Types of pesticides

- A. Insecticides killing plants
- B. Herbicides.....killing insects
- C. Fungicides killing mice
- D. Rodenticides killing rodents and rats

Brief: Insecticide: any toxic substance that is used to kill insects. Such substances are used primarily to control pests that infest cultivated plants or to eliminate disease carrying insects in specific areas.

428. The percentage composition of phosphorus and nitrogen in DAP are _____and_____respectively.

- A. 23.48 % and 21.21 %
- C. 60 % and 40 %
- B. 25.5 % and 21.21 %
- D. 27.9 % and 21.5 %

Brief: Molar mass of $(\text{NH}_4)_2\text{HPO}_4 = 132 \text{ g/mol}$

$$P\% = \frac{31}{132} \times 100 \% = 23.48 \%$$

$$N\% = \frac{28}{132} \times 100\% = 21.21 \%$$

429. What is the purpose of adding coal in cement manufacturing

- A. As energy source in cement production
- B. As dehydrating agent in cement production
- C. It is used to catalyzing the reaction
- D. It is used to hydrolysis reagent.

Brief:

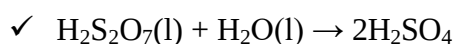
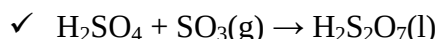
- ✚ Coal is used as an energy source in cement production.
- ✚ Large amounts of energy are required to produce cement.
- ✚ Kilns usually burn coal in the form of powder and consume around 450g of coal for about 900g of cement produced.

430. Which of the following are the correct steps in sulphuric acid (98%)manufacturing process

- A. Sulphur $\rightarrow \text{SO}_2 \rightarrow \text{H}_2\text{S}_2\text{O}_2 \rightarrow \text{H}_2\text{SO}_4$
- B. Sulphur $\rightarrow \text{SO}_2 \rightarrow \text{SO}_3 \rightarrow \text{H}_2\text{S}_2\text{O}_2 \rightarrow \text{H}_2\text{SO}_4$
- C. Sulphur $\rightarrow \text{SO}_2 \rightarrow \text{SO}_3 \rightarrow \text{H}_2\text{S}_2\text{O}_7 \rightarrow \text{H}_2\text{SO}_3$
- D. Sulphur $\rightarrow \text{SO}_2 \rightarrow \text{SO}_3 \rightarrow \text{H}_2\text{S}_2\text{O}_2 \rightarrow \text{H}_2\text{SO}_4$

Brief:

- ✚ Preparation of sulphur dioxide: SO_2 is prepared by burning sulphur in the presence of excess air.
 $\text{S(s)} + \text{O}_2 \text{ (g)} \rightarrow \text{SO}_2 \text{ (g)}$
- ✚ Preparation of sulphur trioxide: Sulphur trioxide is formed when sulphur dioxide reacts with oxygen in a ratio of 1:1 at a temperature of $400^\circ\text{C} - 450^\circ\text{C}$ and a pressure of 1-2 atm in the presence of V_2O_5 as a catalyst. $2\text{SO}_2 \text{ (g)} + \text{O}_2 \text{ (g)} \rightarrow 2\text{SO}_3 \text{ (g)}$
- ✚ Preparation of concentrated sulphuric acid: The sulphur trioxide formed is first made to react with concentrated sulphuric acid. Sulphur trioxide cannot be dissolved in water directly as it leads to the formation of fog. The product obtained after this reaction is known as oleum. The oleum obtained is then dissolved in water to obtain concentrated sulphuric acid.



- ✚ Types of pesticides:
- ✚ Insecticides (killing insects)
- ✚ Herbicides (killing plants)
- ✚ Fungicides (killing fungus)
- ✚ Rodenticides (killing rodents, like mice)
- ✚ Bactericides (killing bacteria)

The most important chemical groups of herbicides are chlorophenoxy acids such as 2,4-D and 2,4,5-T; triazines such as atrazine, hexazinone, and simazine; organic phosphorus chemicals such as glyphosate; amides such as alachlor and metolachlor; thiocarbamates such as butylate; dinitroanilines such as trifluralin; chloroaliphatics such as dalapon and trichloroacetate; and inorganic chemicals such as various arsenicals, cyanates, and chlorates.

431. Which of the following is correct about the industrial manufacturing of sodium carbonate (Na_2CO_3)
- A. The raw materials are: Brine solution (NaCl), Limestone (CaCO_3), Ammonia and Water
 - B. The byproduct of manufacturing of Na_2CO_3 in the Solvay process is CaCl_2
 - C. Burning coke in the Solvay process is required to provide the heat energy
 - D. The Solvay process involves industrial manufacturing process of Na_2SO_4

Brief:

- ✓ Four major raw materials: NaCl , CaCO_3 , NH_3 , H_2O
- ✓ Manufacturing process is Solvay process
- ✓ Coke is used as source of energy to decompose lime stone

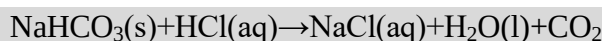
432. Which of the following is NOT the correct chemical formula of the following compounds?

- | | |
|---|--|
| A. Washing soda..... Na_2CO_3 | C. Gypsum..... $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ |
| B. Baking soda..... Mg_2CO_3 | D. Lime water..... $\text{Ca}(\text{OH})_2$ |

433. Which of the following compounds are administered as antacids?
- | | |
|------------------------|-----------------------------|
| A. Sodium carbonate | C. Sodium hydrogencarbonate |
| B. Aluminium carbonate | D. Magnesium acetate |

Brief:

The chemical formula of sodium hydrogen carbonate is NaHCO_3 is often used as an antacid because it neutralizes excess of HCl secreted from the walls of the stomach



However, excessive hydrogen carbonate can make the stomach alkaline and trigger the production of even more acid. Milk of magnesia, which is an aqueous solution of magnesium hydroxide, is also used as an antacid.

434. Given the following statements about Glass, which one is correctly described

- I. Glass is an inorganic solid material that is usually transparent or translucent that is used in numerous applications in our daily lives.
 - II. Glass is made from natural and abundant raw materials such as sand, soda ash, and limestone and that is melted at a very high temperature to form a new material, glass.
 - III. The material required to form soda lime glass is Chalk, soda ash (Na_2CO_3) and sand.
 - IV. The material required to form potash lime of glass is Chalk, potassium carbonate (K_2CO_3), and clean fresh sand.
 - V. The material required to form this type of glass is Chalk, salt cake (Na_2SO_4) coke, and ordinary sand
- A. I,II,III,IV,V B. I,II,V C. II,III,IV D. III,IV,V

Brief:

 The entire above are correctly described about the manufacturing of Glass is

- A. Batch preparation→Glass melting→Glass forming→Annealing→Inspection→Packing
- B. Glass melting→ Batch preparation →Glass forming→Annealing→Inspection→Packing
- C. Glass melting→ Annealing → Batch preparation →Glass forming→Inspection→Packing
- D. Inspection→Glass melting→ Annealing → Batch preparation →Glass forming→Packing

Brief:

➤ Steps in glass production

- I. Batch preparation: the preparation of raw materials is mixed in a proportion of 60% sand, 21% sodium carbonate and 19% limestone.
- II. Glass melting: The raw materials and recycled glass (according to their colour) are fed into a glass-melting furnace.
- III. Glass forming: This is the stage in which the melted glass is made into a required shape.
- IV. Annealing: This is a process that involves the removal of internal stresses by reheating the glass followed by a controlled slow-cooling cycle
- V. Inspection: Testing the quality of glass produced
- VI. Packing and dispatching is the final stage before distribution

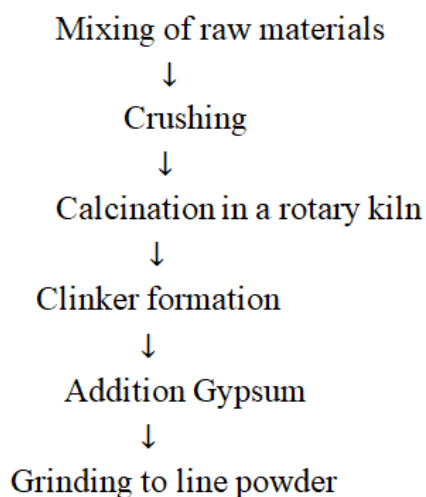
435. For the manufacture of cement various steps are given below. Arrange the steps in a proper sequence.

- | | |
|------------------------|--------------------------------|
| (a)Clinker formation | (d) Mixing of raw materials |
| (b)Crushing | (e) Grinding to fine powder |
| (c) Addition of gypsum | (f) Calcination in rotary kiln |

- A. b d f c e a
B. a b c d e f

- C. b f d c e a
D. d f b e c a

Brief:



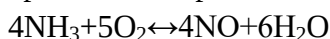
436. The steps for the manufacture of nitric acid by Ostwald process are given. which of the following is the correct sequence.

- a. Recirculation of nitric oxide
- b. Addition of H_2O to nitrogen dioxide
- c. Oxidation of nitric oxide.
- d. Catalytic oxidation of ammonia

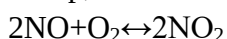
A. d c b a B. dcab c. abcd d, dcba

Brief:

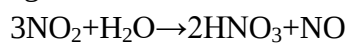
- ❖ In the first step of Ostwald process, ammonia is oxidized to form nitric oxide.



- ❖ In the second step, nitric oxide is oxidised to form nitrogen dioxide.



- ❖ Then, nitrogen dioxide is mixed with water in an absorption tower to get nitric acid:



- ❖ Lastly, nitrogen dioxide is recirculated in the absorption tower to ensure that more and more of it is absorbed with water to increase the concentration of nitric acid.

- ❖ The vapors of nitric acid are then passed over concentrated sulphuric acid that absorbs water from it to increase its concentration further.

439. What is the role of caustic soda in the industrial cleaning process?

- A. To dissolves the grease, protein-based deposits, fats, and oil in manufacturing process
- B. Used as a catalyst
- C. Providing the intermediate of the reaction provided
- D. To increase the dehydration process of chemical reaction

440. For which of the following chemical reaction (process) NaOH is NOT to be important

- A. Saponification B. Etherification C. esterification reactions D. Carboncation reaction

Brief:

- ✓ Sodium hydroxide (NaOH) is a strong base that can act as a catalyst in a variety of reactions by lowering activation and provides higher yield of desired products

- ✓ **Etherification** is the well-known dehydration of an alcohol to form ethers.

- ✓ **Saponification** reaction consumes an important fraction of the catalyst.

- ✓ **Esterification:** is a reaction between an acid and an alcohol to form an ester.

- ✓ **Hydrolysis:** is the reaction of a compound with water to form two or more simpler unite

441. Which of the following are the major raw materials in glass manufacturing process

- A. sand, soda ash and limestone
- B. Borax, potash, sand
- C. Gypsum, sodash, Alumina
- D. Silicate, Gypsum, Sand

Brief:

✚ There are in total four main ingredients that make up glass, these are

✚ Sand (SiO_2)

✚ Limestone (CaCO_3)

✚ Soda Ash (Na_2CO_3)

✚ Recycled Glass.

442. Which one of the following is not used as a raw material in the manufacture of glass?

A. Soda B. Alumina C. Gypsum D. Borax

Brief: Gypsum is not the main raw material in the manufacturing process of glass, but it is used

✚ in cement manufacturing.

443. Consider the following statements: Which of the statements given above is/are correct?

1. Baking soda is used in fire extinguishers.
2. Quicklime is used in the manufacture of glass.
3. Gypsum is used in the manufacture of Plaster of Paris.

A. 1 and 2 B. 2 and 3 C. 1, 2 and 3 D. 1 and 3

Brief:

❖ Uses of Baking Soda

- ✓ Reduces the acidity in the stomach.
- ✓ Acts as an antacid which is used to treat stomach upset and indigestion.
- ✓ Used in the process of washing as a water softener.
- ✓ Due to the formation of soapy foam, it is used in fire extinguishers.

❖ Uses of Quicklime

- ✓ Manufacturing of cement, paper, and high-grade steel.
- ✓ as a reagent in laboratories for dehydration
- ✓ precipitation reaction
- ✓ As ingredient in the manufacturing of caustic

❖ Uses of Gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$)

- ✓ fluxing agent
- ✓ fertilizer
- ✓ filler in paper and textiles
- ✓ Retarder in Portland cement

444. Which of the raw material is Not used for the manufacture of Portland cement

A. Gypsum B. CaCO_3 C. Sand D. CaNO_3

Brief:

The raw materials for the production of cement are

- | | | | | |
|---|---|-------|--|--|
| <ul style="list-style-type: none"> • Limestone (major) • Clay • Gypsum | } | Major | <ul style="list-style-type: none"> • Silica sand • calcium silicate • calcium aluminate | <ul style="list-style-type: none"> • iron (III) oxide • magnesium oxide • pumice. |
|---|---|-------|--|--|

445. Which of the following pairs are mismatched about cement and ceramic

- Gypsum.....setting retarder of cement
- Water Cement plasticizer
- Clinker.....silicates and aluminates
- H₂O₂.....Bleaching agent in sugar manufacturing

Brief:

- ✓ Hydrogen peroxide is the most frequently used chemical for high-yield pulp bleaching when high levels of brightness are required.

446. The process for which cement is made by heating limestone (chalk), alumina (Al₂O₃) and silica-bearing materials such as clay to 1450 °C in a kiln is

- A. calcination B. vulcanization C. Clinker D. Bleaching

447. Which of is the correct difference between calcination and roasting

- In calcination ore is heated in the absence of air while in roasting ore is heated in the presence of oxygen
- In calcination ore is heated in the presence of air while in roasting ore is heated in the absence of air
- In calcination SO₂ is produced while CO₂ is produced in roasting
- calcination is used in sulphide ore while calcination is used in carbonate ore

Brief:

Calcination	Roasting
It is used for carbonates and oxide ores.	It is used for sulphide ores.
Moisture and organic impurities are removed.	Volatile impurities are removed as oxides such as SO ₂ , P ₂ O ₅ , As ₂ O ₄ .

448. Which of the following is true regarding nitric acid?

- A. It is a strong reducing agent
- B. It is a weak oxidizing agent
- C. Its basicity is unity
- D. It is non-planar in gaseous state

Brief:

Nitric acid is a very weak reducing agent since it has a polar O – H bond. This breaks to donate the H^+ ion which is why it is a strong oxidizing agent and a strong acid. Since there is only one cleavable O – H bond, the basicity of nitric acid is unity (one HNO_3 molecule can donate only 1 H^+ ion). It exists as a planar molecule in vapor phase.

449. Which of the following is not an oxi-acid of nitrogen?

- A. Hyponitric acid
- B. Hyponitrous acid
- C. Nitrous acid
- D. Nitric acid

Brief: Oxi-acids are containing oxygen as main component, these are HNO_3 , H_2SO_4 , H_3O_4

450. Which of the following reactions best represents lab scale preparation of nitric acid?

- 1. $3HNO_2 \rightarrow HNO_3 + H_2O + 2NO$
- 2. $NO_2 + O_2 \rightarrow NO_3$
- 3. $O_3 + H_2SO_4 \rightarrow NaHSO_4 + HNO_3$
- 4. $3NO_2 + H_2O \rightarrow 2HNO_3 + NO$

Brief:

The most appropriate lab scale preparation method of nitric acid, HNO_3 is using an alkali nitrate salt and react it with concentration sulphuric acid in a glass retort. Nitrous acid being highly unstable decomposes into nitric acid. The other two sets of reaction represent the industrial process of manufacturing nitric acid i.e. Ostwald's process.

451. What is the name of the industrial process to manufacture nitric acid?

- A. Contact process
- B. Haber-Bosch process
- C. Solvay process
- D. Ostwald's process

Brief:

- ✓ Ostwald's process is the name of the industrial process to manufacture nitric acid in bulk.
- ✓ It involves the oxidation of ammonia which forms nitric oxide. This is then reacted with more oxygen to produce nitrogen dioxide.
- ✓ Subsequently, nitrogen dioxide is dissolved in water to produce adequate concentrations of nitric acid.
- ✓ Contact process is used to produce sulfuric acid.
- ✓ Solvay is used to obtain sodium carbonate
- ✓ Haber-Process to obtain ammonia.

452. What is the catalyst used in the industrial manufacture of nitric acid?

- A. Powdered iron (III) oxide C. Vanadium (V) oxide
B. Zinc-mercury amalgam D. Platinum-Rhodium gauze sheet

Brief:

- ✓ Pt-Rh gauze sheet is widely used as the catalyst in ammoniac oxidation, the first step of Ostwald's process.
- ✓ Fe filling is used in Haber's process
- ✓ V_2O_5 in contact process

453. Which of the following are the most common important components of fertilizer

- A. Nitrogen, phosphorus, sulphur C. Chlorine Bromine, Fluorine
B. Nitrogen, Oxygen, sulphur D. Nitrogen, phosphorus, potassium

Brief:

The three important elements in most fertilizers are

- Nitrogen (N) e.g. from ammonium or nitrate salts like ammonium sulphate, ammonium sulphate or ammonium phosphate (e.g. look for the N in the formula of ammonium salts) or urea, formula $CO(NH_2)_2$.
- Phosphorus (P) e.g. from potassium phosphate or ammonium phosphate.
- Potassium (K) e.g. from potassium phosphate or potassium sulphate.
- The fertiliser is marked with an 'NPK' value, i.e. the nitrogen : phosphorus : potassium ratio.

- Different NPK formulations are made up to suite a particular soil or crop and often applied in slowly dissolving pellet form. Whatever the particular 'NPK' compounds used, the fertilizer components must be soluble in water to be taken in by plant roots.

454. In Habber process, Ammonia gas is collected by-

- B. Upward displacement of water
- C. Upward displacement of air
- D. Downward displacement of water
- E. Downward displacement of air

Brief:

Since ammonia gas is lighter than air, it is collected by downward displacement of air.

455. Ammonia gas is not prepared by:

- A. The direct reaction between Nitrogen and Hydrogen
- B. Hydrolysis of calcium cyanamide.
- C. Heating ammonium chloride with alkaline base.
- D. Heating of NH_4NO_3

Brief:

Heating of Ammonium nitrate is not used in the preparation of ammonia as it is explosive in nature and it decomposes forming nitrous oxide and water vapors.

456. Which of the following conditions is NOT correctly required in the Haber's process-

- A. Temperature should be = 400-500 degree Celsius
- B. Pressure should be = 200-800atm
- C. Catalyst = finely divided iron (Fe).
- D. The Temperature should be at negative level

457. Which of the following is NOT true about Habber process

- A. Both Nitrogen and Hydrogen gases used for preparation of ammonia are sourced from air
- B. During production of ammonia gas, nitrogen and hydrogen gas combines in the ratio of 1:3 by volume.
- C. Adding of more reactants that are nitrogen and hydrogen are important during the reaction in order to regulate the temperature.
- D. Iron felling catalyst used is generally mixed with Al_2O_3 or K_2O which acts as a promoter.

Brief:

- ✚ The ammonia gas is produced by the reaction between nitrogen and hydrogen. It is a reversible and exothermic chemical reaction. Iron is used as a catalyst

- ✚ In the Haber process of ammonia production, iron is used as a catalyst. The catalyst used is generally mixed with Al_2O_3 or K_2O which acts as a promoter. A promoter refers to the substance added to the catalyst to increase its efficiency.
- ✚ The ideal conditions for favouring the forward reaction in ammonia production is high pressure of about 200 atm and high temperature of about 700K. Cooling is very important during this reaction as the reaction is exothermic. Hence, cool reactants that are nitrogen and hydrogen are added in the reactors during the reaction in order to reduce the temperature.
- ✚ The nitrogen used for preparation of ammonia is sourced from air and hydrogen comes from natural gas and water. During production of ammonia gas, nitrogen and hydrogen gas combines in the ratio of 1:3 by volume. These gases are first heated and then passed over a catalyst under high pressure. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

458. Which of the following is the most important chemical in the manufacturing of Ammonia

- | | |
|----------------------|------------------------|
| A. Calcium hydroxide | C. Potassium hydroxide |
| B. Sodium hydroxide | D. Sulphuric acid |

Brief:

Caustic soda (NaOH) is a stronger alkali than quicklime (CaO) and it reacts with glass forming sodium silicate, especially at higher temperatures. Since the production of ammonia requires heating the alkali with an ammonium salt, using the milder alkali of calcium appears to be a better option.

459. How might solid sodium carbonate be obtained from sodium carbonate solution?

- A . Centrifugation B. Filtration C. Evaporation D. It cannot be extracted

Brief:

- ✚ Sodium carbonate obtained from a sodium carbonate solution by evaporates the solvent.

Pure sodium carbonate is a white, crystalline solid.

Sodium carbonate at present is mostly mined from its natural deposits. It also is manufactured synthetically by Solvay (or ammonia-soda) process. The natural production of sodium carbonate currently has surpassed its synthetic production.

The Solvay process involves a series of partial reactions.

The first step is calcination of calcium carbonate to form lime and CO_2 . Lime is converted to calcium hydroxide.

460. Which of the following is method of sodium carbonate preparation

- A. Extracting from the ashes of plants grown in Sodium rich soil.
- B. Solvay process
- C. The reaction between acid and base
- D. Down cell process

Brief:

- ❖ Historically, Sodium Carbonate was extracted from the ashes of plants grown in Sodium rich soil.
- ❖ Presently it is prepared by various methods like, Solvay process, Leblanc process, electrolytic process.
- ❖ Sodium carbonate is synthetically produced in large quantities from salt (sodium chloride) and limestone by a method known as the Solvay process. Sodium carbonate (Na_2CO_3) is the sodium salt of carbonic acid. Sodium carbonate is also known as soda, washing soda, soda ash and soda crystals.
- ❖ Sodium carbonate is formed from the reaction between carbonic acid and sodium hydroxide.

461. The 4 most abundant elements in the earth's crust in decreasing order of abundance are

- A. Oxygen, Silicon, Aluminium and Iron
- B. Aluminium, Iron, Calcium and Magnesium
- C. Iron, Aluminium, Silicon and Oxygen
- D. Silicon, Aluminium, Magnesium and Sodium

Brief:

- ✓ Oxygen is the most abundant element in the Earth's crust that is 46.6% Oxygen by weight.
- ✓ The second most abundant element is Silicon at 27.7% by weight.
- ✓ Other elements in decreasing order of abundance are : Aluminium, Iron, Calcium, Sodium, Potassium and Magnesium
- ✓ Abundance : Oxygen > Silicon > Aluminium > Iron > Calcium > Sodium > Potassium > Magnesium

462. Which of the following **does not** add CO₂ to the atmosphere?

- A. Burning of wood
 - B. Photosynthesis
 - C. Burning of kerosene
 - D. use of motorcars
- The Process of photosynthesis takes up Carbon dioxide and release oxygen.
- Hence photosynthesis does not add carbon dioxide to the carbon cycle.
- Carbon dioxide is added to the atmosphere naturally when organisms respire or decompose (decay), carbonate rocks are weathered, forest fires occur, and volcanoes erupt.
- Carbon dioxide is also added to the atmosphere through human activities, such as the burning of fossil fuels and forests and the production of cement.

463. Which industrial chemicals **mismatched** with its manufacturing process

- A. NH₃ - Haber process C. H₂SO₄-contact process
- B. HNO₃ - Solvay process D. S- Frash process

Brief:

Chemicals	Manufacturing process	
NH ₃	Habber process	Iron filling
HNO ₃	Ostwald process	Platinum and palladium are commonly used as catalyst.
H ₂ SO ₄	Contact process	Vanadium pent oxide (V ₂ O ₅) is catalyst
Na	Down cell process	
Al	Hall process	Cryolite (Na ₃ AlF ₆) is used for reducing the melting point.
	Frash process	The three raw materials are:-
Sulphur		1. Plant/tree-based – materials 2. Animal-based– materials 3. Mining-based– materials

464. Which of the following is a natural polymer?

- A. PVC B. Polyethene C. Cellulose D. Teflon

465. Which of the following statements is Not correctly described about a given of each concept

- A. Ethanol is used in this experiment to dissolve oil or fat and make hydrolysis occur easily.
- B. Sodium chloride solution is added to reduce the solubility of soap and separate it out from the solution
- C. Detergent lathered well or easily form lather with hard water than soap does.
- D. The formation of leather is indicating the less solubility of soap in water.

466. Given the following properties/uses/ of sulphuric acid, which is/are correct?

- I. Sulphuric acid is used in the production of fertilizers, detergents, and explosives
- II. Sulphuric acid is an oxidizing agent and a good dehydrating agent.
- III. It is used to dry neutral and acidic gases such as nitrogen, oxygen, and carbon dioxide.
- IV. It is used as a catalyst in the manufacture of many chemicals.

- A. I, II, III, IV B. I, II, IV C. II and IV D. IV only

UNIT TEN

10. ENVIRONMENTAL CHEMISTRY

Unit outcome

- *At the end of this unit, you will be able to*
 - ❖ *Define environmental chemistry, pollution, pollutants*
 - ❖ *Describe the components that make the environment*
 - ❖ *Explain environmental pollution*
 - ❖ *List source of pollutants, effect and their mitigation for water pollution, air and land pollution*
 - ❖ *Discuss the causes and effects of air, water and land pollution*
 - ❖ *Define terms: Greenhouse gas, Global warming, green chemistry*
 - ❖ *List greenhouse gases and describe their level of effect*
 - ❖ *Discuss about global warming and climate change*
 - ❖ *State about pollution control mechanisms for controlling, air, water and land pollution*

10.1 .Introduction

Environmental chemistry is the scientific study of the chemical and biochemical phenomena occurring in natural environments. It focuses on understanding the sources, reactions, transport, effects, and fates of chemical species in air, water, and soil. It also examines the impact of human activities on the environment and explores ways to mitigate pollution and promote sustainability.

10.1.1. Scope of Environmental Chemistry

1. Atmospheric Chemistry: Study of the chemical composition of the atmosphere, air pollutants, and their effects.
2. Aquatic Chemistry: Investigation of chemical interactions in natural water bodies, including lakes, rivers, and oceans.
3. Soil Chemistry: Analysis of soil components, nutrients, and pollutants.
4. Toxicology and Ecotoxicology: Examination of the effects of chemicals on human health and ecosystems.
5. Waste Management and Remediation: Methods to control, treat, and remediate pollution.
6. Green Chemistry: Development of sustainable and environmentally friendly chemical processes.

10.1.2. Air Pollution and Atmospheric Chemistry

- Sources of Air Pollution: Industrial emissions, vehicle exhaust, natural sources (volcanic eruptions, wildfires).
- Major Air Pollutants:
 - Carbon monoxide (CO)
 - Sulfur dioxide (SO₂)
 - Nitrogen oxides (NO_x)
 - Particulate matter (PM₁₀, PM_{2.5})
 - Ozone (O₃)
 - Volatile organic compounds (VOCs)

❖ Effects of Air Pollution:

- ✓ Respiratory diseases (asthma, lung cancer)
- ✓ Acid rain (caused by SO₂ and NO_x)
- ✓ Ozone layer depletion (due to CFCs and halons)
- ✓ Greenhouse effect and global warming

10.1.3. Water Pollution and Aquatic Chemistry

❖ Sources of Water Pollution:

- Industrial waste discharge
- Agricultural runoff (pesticides, fertilizers)
- Domestic sewage
- Oil spills

❖ Types of Water Pollutants:

- Heavy metals (lead, mercury, cadmium)
- Organic pollutants (pesticides, detergents, pharmaceuticals)
- Pathogens (bacteria, viruses)

❖ Effects of Water Pollution:

- Eutrophication (excess nutrients leading to algal blooms)
- Bioaccumulation and biomagnification (toxins building up in organisms)
- Drinking water contamination

10.1.4. Soil Pollution and Land Degradation

❖ Causes of Soil Pollution:

- Use of chemical fertilizers and pesticides
- Industrial waste dumping
- Deforestation and land misuse

❖ Major Soil Pollutants:

- Heavy metals (arsenic, lead, mercury)
- Organic pollutants (PAHs, dioxins)
- Radioactive substances

❖ Effects of Soil Pollution:

- Loss of soil fertility
- Groundwater contamination
- Harmful effects on plant and animal life

10.1.5. Environmental Chemistry and Climate Change

❖ Greenhouse Gases (GHGs):

- Carbon dioxide (CO₂)
- Methane (CH₄)
- Nitrous oxide (N₂O)
- Fluorinated gases (CFCs, HFCs)

❖ Effects of Climate Change:

- Rising global temperatures
- Melting polar ice caps
- Rising sea levels
- Extreme weather events (hurricanes, floods, droughts)

❖ Mitigation Strategies:

- Renewable energy sources (solar, wind, hydroelectric)
- Carbon capture and storage (CCS)
- Sustainable agriculture
- Reduction of fossil fuel consumption

10.2. Questions with Explanation

467. Which of the following is the most abundant greenhouse gas in the atmosphere

- A. carbon dioxide B. Carbon monoxide C. water vapour D. sulphur dioxide

Brief:

Water vapour is the most abundant greenhouse gas. It increases as the earth's atmosphere warms but unlike CO₂, which can remain in the earth's atmosphere for centuries, water vapour persists for a few days. But, Carbon dioxide is widely reported as the most important anthropogenic greenhouse gas because it currently accounts for the greatest portion of the warming associated with human activities. Oxygen is not a greenhouse gas. Carbon monoxide is a very weak greenhouse gas, but has quite a few indirect effects on global warming.

468. The gases that contribute to the greenhouse effect on Earth are, in order of greatest to smallest in importance,

A. CO₂, H₂O, CH₄

C. CH₄, CO₂, H₂O

B. H₂O, CO₂, CH₄

D. H₂O, CH₄, CO₂

Brief:

- water vapour (H₂O)
- carbon dioxide (CO₂)
- nitrous oxide (N₂O)
- methane (CH₄)
- ozone (O₃)

1: The sun emits visible radiation.

2: Visible radiation passes through the Earth's atmosphere.

3: The ground absorbs visible radiation and emits infrared radiation.

4: Greenhouse gases absorb and emit infrared

469. On average, how do low clouds affect the surface temperature (compared to a clear sky)?
- A. Low clouds warm the surface by increasing the greenhouse effect
 - B. Low clouds cool the surface by increasing the energy emitted to space
 - C. Low clouds cool the surface by increasing the fraction of light that a surface reflects
 - D. It depends on whether they are comprised of ice or liquid water

470. Which One of the Following Is Not a Greenhouse Gas?

- A. Methane B. Hydrogen C. Nitrous oxide D. Ozone

Brief:

Methane, carbon dioxide, nitrous oxide and water vapour are the primary greenhouse gases that contribute to the greenhouse effect. A greenhouse gas (GHG or GhG) is a gas that absorbs and emits radiant energy in the thermal infrared spectrum, resulting in the greenhouse effect.

471. Normal greenhouse effect is important for the sustenance of life on earth as it has increased the earth's surface temperature by

- A. 15 °C B. 50 °C C. – 18 °C D. 33 °C

Brief:

The greenhouse effect is a natural phenomenon that keeps the earth at the temperature required for life to thrive. The remaining radiation, roughly half of it, reaches the planet and is absorbed by seas and land. This keeps the ground warm and provides food for plants, animals, and humans increasing of temperature by 33 °C.

472. Which of the Following Greenhouse Gases Is Present in Very High Quantities?

- A. Carbon dioxide B. Ethane C. Propane D. Methane

Brief:

Carbon dioxide has the highest concentration among the greenhouse gases, the concentration ranges from 9 to 26%, another important gas methane constitutes about 4 to 9% whereas ozone constitutes about 3 to 7%

475. The Most Potent Greenhouse Gas in Terms of Efficiency Is

- A. N_2O B. C_2O C. CFC D. CH_4

Brief:

CFC stands for chloro-fluoro carbons, these gases are primarily responsible for the ozone layer depletion, and additionally they also trap heat in the atmosphere thus contributing to the greenhouse effect.

476. Which One of the Following Is Incorrect About the Greenhouse Effect?

- A. Life on earth is possible due to the greenhouse effect
B. Greenhouse effect is a natural process that maintains the earth's temperature
C. More is the emission of greenhouse gases, more is the temperature of the earth's atmosphere
D. Increased emission of greenhouse gases is a natural process

Brief:

Increased emission of greenhouse gases is not a natural process, it rather occurs due to increased industrial activities.

477. Which is Not Naturally Occurring Greenhouse Gas?

- A. Nitrous oxide B. Methane C. Carbon dioxide D. Ethane

Brief:

Ethane is found at very low concentrations thus does not significantly contribute to the greenhouse effects, rather the conversion of ethane to methane largely contributes to the greenhouse effects.

478. Which is Not True for the Relative Contribution of Various Greenhouse Gases to Total Global Warming?

- | | |
|-------------------------|-------------------------------|
| A. CFCs – 14% | C. N_2O – 12% |
| B. Carbon dioxide – 60% | D. Methane – 20% |

Brief:

N_2O - 12%, All the mentioned gases are contributors to global warming, but the relative global warming potential of N_2O is 6 percent. Relative global warming potential is the measurement of the amount of heat trapped in the atmosphere by a greenhouse gas over a specific period of time relative to CO_2 .

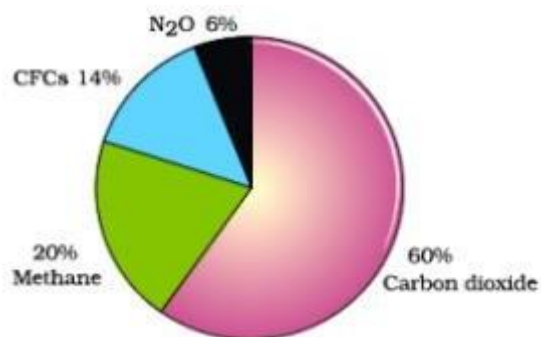
479. Arrange CFC, CH₄, N₂O and CO₂ in decreasing order according to their contribution in global warming.

A. CO₂ > N₂O > CFC > CH₄

C. CH₄ > CFC > N₂O > CO₂

B. CFC > CO₂ > CH₄ > N₂O

D. CO₂ > CH₄ > CFC > N₂O



The diagrammatic representation of relative contribution of various greenhouse gases to total global warming is shown above. From the chart, it is observed that their contribution in global warming in decreasing order is: CO₂ > CH₄ > CFC > N₂O.

480. The GWP (The global warming potential) is dependent on which of the following factors?

A. Absorption capability of Infra-Red radiations

B. Atmospheric lifetime

C. Range of Infrared wavelengths it can absorb

D. All of the above

Brief:

The heat absorbed by any greenhouse gas in the atmosphere as a multiple of the heat absorbed by the same quantity of carbon dioxide is referred to as global warming potential. All the factors are contributors to the GWP.

481. What is relative contribution of H₂O, CO₂, CH₄, CFC and N₂O to global warming?

A. 20%, 14%, 60%, 6%, 85%,

C. 85%, 60%, 20%, 14%, 6%

B. 14%, 6%, 20%, 85%, 60%

D. 6%, 14%, 20%, 60%

Brief:

- ❖ Global warming potential (GWP) is a measure of how much heat a greenhouse gas traps in the atmosphere up to a specific time horizon, relative to carbon dioxide.
- ❖ The major contribution is from CO₂ (60%),CH₄ (20%), CFC and N₂O attributes to 14% and 6% respectively. So the correct answer is '60%,20%,14%,6%'.

482. Anthropogenic Sources of Carbon Dioxide Are 20 Times More Than Natural Carbon Dioxide Sources.

A.True B. False

Brief:

CO₂ emissions from power generation, transportation, industrial sources, chemical manufacture, petroleum production, and agricultural operations are all examples of anthropogenic CO₂ emissions. Many of these sources emit CO₂ as a consequence of burning fossil fuels (coal, oil, and natural gas).

483. What is Carbon Sequestration?

- A. Removal of CO₂ from the atmosphere
- B. Storage of CO₂ by depositing in the reservoir
- C. Removal of CO₂ from the atmosphere & storing it by depositing it in the reservoir
- D. Removal of CO from the atmosphere & storing it by depositing it in the reservoir

Brief:

The method of trapping and storing atmospheric carbon dioxide is known as carbon sequestration. It is one way of lowering carbon dioxide levels in the atmosphere with the purpose of slowing global warming.

484. One of the following is NOT Naturally Occurring Greenhouse Gas?

- A. Nitrous oxide B. Methane C. Carbon dioxide D. Ethane

Brief:

There are both natural and human-caused greenhouse gases. Natural sources include respiration and decomposition of plants and ocean release of greenhouse gases to the atmosphere. Many natural GHGs occur naturally in the atmosphere, such as water vapour, carbon dioxide, methane and nitrous oxide.

485. What are the three R's that are used to save the environment which includes forest?

A. Reduce, Reuse, Recycle

C. Reuse, Reserve, Reduce

B. Reserve, Reduce,
Recycle

D. Reserve, Reuse, Reduce

Brief:

The three R's reducing, reuse and recycle - are three approaches, and the most environmentally preferred. Reducing, reusing and recycling waste helps save landfill space by keeping useful materials out.

486. Which of the following statements are correct about the atoms economy method, EXCEPT?

A. The atoms economy analysis encourages industry to think about finding ways to have the greatest percentage of atoms end up in an actual product.

B. The atoms economy helps us determine the economics of any particular reaction by encouraging waste minimization.

C. Given the reaction (unbalanced) $\text{CO}_2 + \text{H}_2 \rightarrow \text{C}_2\text{H}_5\text{OH} + \text{H}_2\text{O}$ the atoms economy for the production of ethanol is equal with its molecular mass.

D. The greater the % atom economy of a reaction, the more 'efficient' of waste management to be likely.

E. Non-renewable resources are desired for atom economy than renewable raw materials

Brief:

❖ The atom economy (a measure of atom utilization or efficiency) is a measure of the amount of starting materials that end up as useful products.

❖ To develop processes based on renewable (plan-based) rather than nonrenewable (fossil carbon –derived) raw materials.

❖ Good atom economy means most of the atoms of the reactants are incorporated in the desired products and only small amounts of unwanted byproducts are formed and hence lesser problems of waste disposal or waste treatment.

❖ It is calculated by the formula

$$\text{Atom economy} = 100 * \frac{\text{Mass of desired or useful product}}{\text{Total mass of all reactants or products}}$$

E.g. The atom economy ethanol in the reaction: $2\text{CO}_2 + 6\text{H}_2 \rightarrow \text{C}_2\text{H}_5\text{OH} + 3\text{H}_2\text{O}$

$$\text{Atom Economy} = \frac{\text{total mass of ethanol}}{\text{total mass of product}} \times 100\%$$

- Mass of ethanol = 46g/mol
- Mass of H_2O = $3 \times 18\text{g/mol} = 54\text{g/mol}$
- Total mass of product = 100g/mol
- Atom Economy = $\frac{\text{total mass of ethanol}}{\text{total mass of product}} \times 100\%$
- Atom Economy = $\frac{46\text{g/mol}}{100\text{g/mol}} \times 100\% = 46\%$

487. Which of the following is the true about greenhouse gases, greenhouse effect, UV, IR, Global warming and Atmosphere?

- I. Atmosphere → stabilizing earth Temperature and allows the right amount of sunlight in to warm the Earth
- II. Greenhouse gases → water vapor, carbon dioxide, methane and nitrous oxide
- III. Greenhouse effect → layer keeps —not too hot in the summer|| / —not too cold in winter.||
- IV. UV → The solar energy that warms the Earth
- V. IR → The solar radiation that bounces off the Earth back toward the atmosphere
- VI. Global warming → Excess amount gasses in the atmosphere may block heat from escaping into space and trap too much heat

A. I, II, IV, V

C. ALL EXCEPT IV

B. B. II, III, V

D. ALL (I-VI) are correctly matched.

Brief:

- Atmosphere stabilizing earth the temperature and allows the right amount of sunlight in to warm the Earth
- Greenhouse gases → water vapor, carbon dioxide, methane and nitrous oxide
- Greenhouse effect.....layer keeps us —not too hot in the summer|| and —not too cold in the Winter.
- VIS, IR and UV radiation are the solar energy that warms the Earth.

- The solar radiation that bounces off the Earth back toward the atmosphere is in the form of IR with long wave length.
- The main radiation energy that comes from the sun is in the form of Visible light.
- Too many greenhouse gasses in the atmosphere may block heat from escaping into space and trap too much heat and prone the atmosphere in global warming.

488. Things you can do to help decrease global warming include:

- A. Keep your air conditioner on high.
- B. Turn off your lights when you're not using them.
- C. Keep your freezer door open.
- D. Eat more ice cream.

Brief:

- ✚ Greenhouse Gas Emissions are on the rise from air conditioning, and they currently account for almost 4% of the world's overall greenhouse gas emissions.
- ✚ Refrigerants are chemicals used in air conditioning those can cause global warming because they have the potential to trap heat in the Earth's atmosphere on released into the atmosphere.
- ✚ Light pollution, or artificial light at night, is the excessive or poor use of artificial outdoor light, and it disrupts the natural patterns of wildlife, contributes to the increase in carbon dioxide (CO₂) in the atmosphere.
- ✚ Lighting contributes to about 6% of global carbon dioxide emissions. Switching to LED (light emitting diode) lighting could keep tons of greenhouse gases out of the air and minimize the number of power stations.
- ✚ Most gelato machines use Hydrofluorocarbon (HFC) refrigerants that contribute to global warming
- ✚ the global warming potential of chocolate ice cream is very sensitive to land use change associated with cocoa beans cultivation, increasing the impact

489. The energy that warms Earth's surface comes primarily in the form of _____.

- A. heat from Earth's interior
- B. infrared light from the Sun
- C. ultraviolet light from the Sun
- D. visible light from the Sun

Brief:

Earth's temperature remains fairly steady, which means that Earth must return nearly the same amount of energy to space that it receives from the Sun.

The visible light is known to be a portion of the electromagnetic spectrum that can be seen with naked eyes.

Other forms of the electromagnetic radiation are the infrared and ultraviolet.

The visible light is found or falls within the range of electromagnetic radiation spectrum between the infrared and ultraviolet.

Among the electromagnetic radiation, the energy that warms Earth's surface comes primarily in the form of visible light from the Sun.

490. Earth's temperature remains fairly steady, which means that Earth must return nearly the same amount of energy to space that it receives from the Sun. In what form(s) does Earth return most of this energy to space?

- A. ultraviolet light reflected by the surface
- B. visible light reflected by cloud
- C. infrared light emitted by the surface and atmosphere
- D. visible light reflected by the surface
- E. visible light emitted by the surface and atmosphere

Brief:

- * The Earth returns most of the energy it receives from the Sun to space in the form of infrared light emitted by the surface and atmosphere, with a small fraction being reflected as visible light.
- * Most solar energy is absorbed at the surface, while most heat is radiated back to space by the atmosphere. Earth returns most of the energy it receives from the Sun to space in the form of infrared radiation (thermal radiation).

The Sun emits visible light and shortwave radiation, which is absorbed by Earth's surface and atmosphere.

-Earth's surface heats up and, in turn, emit long wave infrared radiation.

-Some of this infrared radiation is absorbed and re-emitted by greenhouse gases in the atmosphere, while the rest escapes directly to space.

-A small amount of energy is also lost through reflected sunlight (visible and shortwave radiation) from clouds, ice, and other reflective surfaces (albedo effect).

491. Greenhouse gases in the atmosphere, such as carbon dioxide and water vapor, make Earth warmer than it would be otherwise because these gases

- A. absorb visible light coming from the Surface
- B. form clouds that emit thermal radiation
- C. reflect visible light coming from the Sun
- D. absorb infrared light emitted by the surface

Brief:

- ✚ While greenhouse gases play a role in maintaining a habitable global temperature, human activity has increased these gases' atmospheric concentration, causing too much heat to become trapped in our atmosphere and warming of earth
- ✚ The extra heat in our atmosphere leads to increased temperatures around the globe and climate change.
- ✚ CO₂ and H₂O absorb heat while CFC causes for global warming by ozone layer depletion.
- ✚ scientists say that doubling pre-industrial carbon dioxide levels will likely cause global average surface temperature to rise between 1.5° and 4.5° Celsius compared to pre-industrial temperatures.

492. According to scientists, the naturally occurring greenhouse effect makes Earth about 31°C warmer than it would be if there were no greenhouse gases in our atmosphere. How do scientists "know" what Earth's temperature would be without greenhouse gases?

- A. They calculate this temperature from Earth's reflectivity and distance from the Sun.
- B. They estimate it by averaging guesses made by many individual scientists.
- C. Ancient fossils allow them to infer Earth's temperature at a time before our atmosphere contained greenhouse gases.
- D. They assume that this temperature would be about the same as the temperature of Mars, which has much less of an atmosphere than Earth.

Brief:

- ✓ Scientist know the temperature of earth without greenhouse gas by studying indirect clues like; the chemical and structural signatures of rocks, fossils, and crystals, ocean sediments, fossil, tree rings and ice core.
- ✓ E.g. Ice cores show that Earth naturally has cooler and warmer periods.

493. Which of the following correctly lists two key pieces of evidence that, together, indicate that we should expect human activity to cause global warming?

- A. (1) The burning of fossil fuels is a form of human activity.(2) Human activity is always bad for the environment.
- B. (1) We can learn about past climate from ice cores.(2) Ice cores show that Earth naturally has cooler and warmer periods.
- C. (1) Carbon dioxide tends to make planets warmer than they would be otherwise.(2) Measurements demonstrate that human activity is raising the carbon dioxide concentration in the atmosphere.
- D. (1) The carbon dioxide cycle increase Earth's climate.(2) Human activity has disrupted the carbon dioxide cycle.

Brief:

- Burning fossil fuels, cutting down forests and farming livestock are increasingly influencing the climate and the earth's temperature by adding of greenhouse gases to those naturally occurring in the atmosphere, increasing the greenhouse effect and global warming.
- Carbon dioxide tends to make planets warmer due to human activity is raising the carbon dioxide concentration in the atmosphere.
- The carbon cycle will keep the concentrations of carbon dioxide in the atmosphere balanced and global temperatures steady.

494. Which of the following catalysts is used in the contact process to sulphuric acid?

- A. Iron
- B. V_2O_5
- C. Vanadium trioxide
- D. Molybdenum

Brief: Vanadium pentoxide is the catalyst used in the contact process.

495. Which of the following is the key step in the manufacture of H_2SO_4 by contact process?

- A. Burning of sulphur of sulphide ores in air.
- B. Absorption of SO_3 in H_2SO_4 to give oleum.
- C. Conversion of SO_2 to SO_3
- D. All of the above

Brief:

- * The key step in the manufacture of H_2SO_4 is the catalytic oxidation of SO_2 with O_2 to give SO_3 in the presence of V_2O_5 (catalyst).

496. Which of the following impurities are usually found during the production of SO_2 in the contact process?

- | | |
|-----------------------|-----------------------|
| A. Hydrogen compounds | C. Nitrogen compounds |
| B. Arsenic compounds | D. Water |

Brief:

- ❖ Purification of the SO_2 produced includes the removal of dust and other impurities such as arsenic compounds.

496. What is the purity percentage of sulphuric acid obtained by contact process?

- | | | | |
|-----------|------------|-----------|-----------|
| A. 96-98% | B. 94-95 % | C. 90-91% | D. 96-98% |
|-----------|------------|-----------|-----------|

Brief

- The Contact process yields 96-98% pure sulphuric acid.
- In the **Contact Process**, sulfuric acid (H_2SO_4) is produced by oxidizing sulfur dioxide (SO_2) to sulfur trioxide (SO_3) using a vanadium pentoxide (V_2O_5) catalyst. The SO_3 is then absorbed in concentrated H_2SO_4 to form **oleum** ($\text{H}_2\text{S}_2\text{O}_7$), which is later diluted with water to obtain concentrated sulfuric acid.
- The final product of the Contact Process is typically **96-98% pure sulfuric acid**, as higher concentrations (like 100%) tend to be unstable due to their high affinity for water.

UNIT ELEVEN

11. POLYMERS

Unit out comes



At the end of this unit, students will be able to

- ❖ *Define terms ; polymer, monomer, polymerization, addition polymerization, condensation polymerization*
- ❖ *Classify polymers based on their sources, whether they are natural or synthetic*
- ❖ *Differentiate Natural polymer and synthetic polymer*
- ❖ *list and explain a variety of synthetic polymers and natural polymers*
- ❖ *classify polymers based on their response to heat*
- ❖ *Developing the skill for predict monomers of the polymers by observing their structures*
- ❖ *Distinguish addition and condensation polymerization*
- ❖ *Developing the skill of writing structure of a polymer given structure of a monomer/monomers*
- ❖ *Developing skill of classing polymers into their types based on different criterion*
- ❖ *Describe the properties of the different types of polymers*
- ❖ *Describe the monomers, properties and uses of plastics, rubbers, carbohydrates and proteins*
- ❖ *Differentiate polymers based on their*

11.1. Introduction

Polymers are large molecules composed of repeating structural units called monomers. These monomers are linked through covalent bonds, forming long-chain structures. Polymers play a crucial role in various industries, including plastics, textiles, and medicine.

11.1.1. Classification of Polymers

✓ Polymers can be classified based on their origin and synthesis process:

1. Natural Polymers – Found in nature. Examples:

- Proteins (e.g., collagen, keratin)
- Carbohydrates (e.g., starch, cellulose)
- Natural rubber

2. Synthetic Polymers – Man-made polymers used in industry. Examples:

- Plastics (e.g., polyethylene, PVC)
- Fibers (e.g., nylon, polyester)
- Elastomers (e.g., neoprene, silicone rubber)

❖ Monomer

- A monomer is a small molecule that acts as the building block for polymers. When monomers link together through polymerization, they form long polymer chains.
- Examples of Monomers and Their Polymers

Monomer	Polymer	Example Use
Ethene (C_2H_4)	Polyethylene (PE)	Plastic bags, bottles
Propene (C_3H_6)	Polypropylene (PP)	Packaging, containers
Vinyl chloride (C_2H_3Cl)	Polyvinyl chloride (PVC)	Pipes, cables
Glucose ($C_6H_{12}O_6$)	Cellulose	Paper, cotton
Terephthalic acid + Ethylene glycol	Polyester	Clothing, bottles

11.1.2. Polymerization

✓ Polymerization is the process of combining monomers to form polymers. It is classified into two types:

* Addition Polymerization

✓ In this type of polymerization, monomers with double or triple bonds react to form a polymer without the loss of any molecules. It occurs in three steps:

- Initiation – A free radical or catalyst starts the reaction.
- Propagation – Monomers add one by one to the growing chain.
- Termination – The reaction stops when two free radicals combine.

Examples of Addition Polymers

Polymer	Monomer	Example Use
Polyethylene (PE)	Ethene	Plastic bags, bottles
Polystyrene (PS)	Styrene	Packaging, insulation
Polyvinyl chloride (PVC)	Vinyl chloride	Pipes, cables
Polytetrafluoroethylene	Tetrafluoroethylene	Non-stick cookware

❖ Condensation Polymerization

This type of polymerization involves the reaction of monomers with the elimination of small molecules such as water (H_2O), ammonia (NH_3), or methanol (CH_3OH).

Examples of Condensation Polymers

Polymer	Monomers	Example Use
Nylon (Polyamide)	Hexamethylenediamine + Adipic acid	Textiles, ropes
Polyester (PET)	Terephthalic acid + Ethylene glycol	Clothing, bottles
Bakelite	Phenol + Formaldehyde	Electrical switches

❖ Thermosetting vs. Thermoplastics

- * Polymers can also be classified based on their behavior when heated.

Thermosetting Plastics

It undergoes permanent hardening upon heating and cannot be remelted due to strong cross-linking.

* Properties of Thermosetting Plastics

- Hard, rigid, and brittle.
- Cannot be reshaped after molding.
- Heat and chemical-resistant.

Examples of Thermosetting Plastics

Polymer	Use
Bakelite	Electrical switches, sockets
Melamine	Kitchenware, heat-resistant plates
Epoxy Resin	Adhesives, coatings

Thermoplastics

These plastics soften when heated and harden when cooled, allowing them to be remolded multiple times.

* Properties of Thermoplastics

- Flexible and lightweight.
- Can be reshaped multiple times.
- Less heat-resistant compared to thermosetting plastics.

Examples of Thermoplastics

Polymer	Use
Polyethylene (PE)	Plastic bags, packaging
Polypropylene (PP)	Containers, automotive parts
Polyvinyl chloride (PVC)	Pipes, flooring
Polystyrene (PS)	Disposable cups, insulation

11.2. Questions with Explanation

497. Which of the following is the chemical name for Dacron?

- | | |
|-------------------------------|-----------------------|
| A. Polyethylene terephthalate | C. Polyvinyl chloride |
| B. Polypropylene | D. Polystyrene |

Brief:

- ❖ Dacron or terylene (a polyester) is formed by the condensation polymerization of **ethylene glycol and terephthalic acid** with the elimination of water molecule.
- ❖ The reaction takes place at 420 - 460 K temperature in presence of zinc acetate and antimony trioxide as a catalyst.
- ❖ Dacron is used to make clothes.

498. What is the monomer unit of nylon?

- | | |
|--------------|--|
| A. Ethylene | C. Adipic acid and hexamethylenediamine |
| B. Propylene | D. Terephthalic acid and ethylene glycol |

Brief:

- ❖ The monomeric repeating unit of Nylon-6 polymer is **NH(CH₂)₅CO** which is derived from caprolactam.
- ❖ The monomeric repeat unit of Nylon-6,6 polymer is **[NH(CH₂)₆NH-CO(CH₂)₄CO]** which is derived from the two monomers,
 - hexamethylene diamine, NH₂(CH₂)₆NH₂ and
 - adipic acid, HO-C(O)(CH₂)₄COOH

499. Which of the following polymers is known for its elasticity and flexibility?

- | | | | |
|----------|-----------|------------|-------------------|
| A. Nylon | B. Dacron | C. Perspex | D. Natural rubber |
|----------|-----------|------------|-------------------|

Brief:

- ❖ Properties of natural rubber

✓ Excellent elasticity	✓ Good wear resistance
✓ Good tensile strength	✓ Excellent cold resistance
✓ High elongation	✓ Good electrical insulator
✓ Good tear resistance	✓ Due to the non-polar high resistance to water.

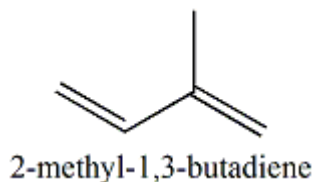
500. What is the monomer unit of natural rubber?

- A. Isoprene
- B. Propylene
- C. Ethylene
- D. Butadiene

Brief:

❖ Natural rubber is polymerized from the monomer isoprene(2-methyl-1,3-butadiene)

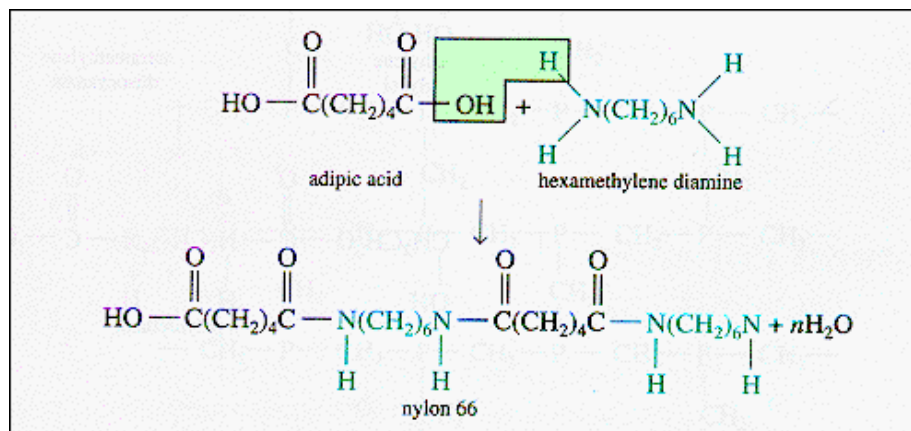
Which is a conjugated diene hydrocarbon?



501. What is the chemical name for nylon?

- A. Polyethylene terephthalate
- B. Polymethyl methacrylate
- C. Polyvinyl chloride
- D. Polyamide

Nylon was the first synthetic thermoplastic polymer; the condensation reaction for making nylons is the reaction of a diamine and a diacid. The product of the reaction is a nylon or polyamide.



Properties of Nylon

- ✓ Lustrous.
- ✓ Elastic.
- ✓ Very strong.
- ✓ Resilient.
- ✓ Does not absorb water.
- ✓ Dries quickly.

502. Which of the following is not a characteristic of condensation polymerization?

- A. Formation of small molecules as byproducts
- B. Loss of water or alcohol during polymerizing rxn
- C. Formation of long chains of repeating units
- D. High temperature and pressure required for polymerization

Brief:

- Common characteristics of condensation polymerization
- ✓ High temperature and pressure required for polymerization
- ✓ The primary structural feature necessary for a molecule to make it useful in a condensation polymerization reaction is the presence of two bifunctional monomer molecules undergoes condensation with the loss of simple molecules of water, alcohol to form a dimer.
- ✓ Condensation polymers form more slowly than addition polymers, often requiring heat, and they are generally lower in molecular weight.
- ✓ The terminal functional groups on a chain remain active, so that groups of shorter chains combine into longer chains in the late stages of polymerization.

503. What is the chemical structure of Perspex?

- | | |
|-------------------------------|----------------------------|
| A. Polyvinyl chloride | C. Polymethyl methacrylate |
| B. Polyethylene terephthalate | D. Polystyrene |

Brief:

- ❖ Perspex is Polymethyl methacrylate (PMMA) which is formed by the free radical polymerization of monomer ***methyl methacrylate***.
- ❖ The polymethyl methacrylate structure is a vinyl polymer which is an ester of methacrylic acid ($\text{CH}_2=\text{C}[\text{CH}_3]\text{CO}_2\text{H}$).
- ❖ **Properties:** Excellent optical clarity & transparency. Highly resistant to variations in temperature. During condensation polymerization, the polymer chains grow by:
 - A. Addition of monomers to the chain
 - B. Elimination of small molecules
 - C. Breaking of covalent bonds
 - D. Formation of ionic bonds

Brief: Condensation polymerization, the polymer chains grow by elimination of small molecules

504. Which of the following conditions is commonly required for condensation polymerization?

- | | |
|---------------------------|----------------------------|
| A. Low temperature | C. High pressure |
| B. Presence of a catalyst | D. Absence of any catalyst |

Brief: Condensation polymerization is facilitated in the presence of different catalysts

505. Which of the following is not a step in addition polymerization?

- | | |
|----------------------|----------------------|
| A. Chain initiation | C. Chain propagation |
| B. Chain termination | D. Chain branching |

Brief:

There are three steps in this type of polymerization:

- ❖ Initiation
- ❖ Propagation
- ❖ Termination.

During initiation, the monomer acquires an active site to become a free radical.

The addition of initiators or other approaches such as absorption of heat, light or irradiation can trigger (initiate) the initiation process.

506. Which of the following is an example of a condensation polymerization reaction?

- A. Formation of polyethylene from ethylene monomers
- B. Formation of polypropylene from propylene monomers
- C. Formation of polystyrene from styrene monomer
- D. Formation of polyester from dicarboxylic acid and diol monomers

Brief:

- ❖ Formation of polyester from dicarboxylic acid and diol monomers is the condensation polymerization reaction

507. What are plasticizers

- A. They are part of monomer
- B. They are catalyst
- C. A substance which is added for strength of polymer
- D. A substance which is added used as catalyst

Brief:

- ❖ A plasticizer is a substance, when added to a material, usually a plastic, makes it flexible, resilient and easier to handle. Early examples of plasticizers include water to soften clay. Today, modern plasticizers are manmade organic chemicals, the majority of which are esters, such as adipates and phthalates.

508. Why are olefins (alkenes) good monomers for addition polymerization reactions?

- A. The electrons in the weak pi-bonds can be used to form strong sigma bonds to other monomer units
- B. Alkenes are easily available raw materials
- C. They are more stable than alkanes
- D. They have high density than alkanes

Brief:

- ❖ Alkenes are hydrocarbons containing one or more carbon-carbon double bonds.

The presence of these double bonds makes alkenes highly reactive, as they are capable of accepting electrons from other molecules

 Why do different polymers have different properties?

- ✓ They have different chemical compositions (different monomer units), different structures, different ways of being fabricated, etc.

 What kind of structural changes accompany bond-breaking and bond-forming in olefin polymerization?

- ✓ The olefin monomers are flat (two-dimensional) molecules with sp^2 - hybridized carbon atoms. The polymers are three-dimensional molecules in which the carbon atoms are sp^3 hybridized.

 What is the role of a catalyst?

- ✓ A catalyst reduces the energy of activation for a reaction by providing an alternative pathway. In this way, it speeds up the reaction and allows it to proceed under milder conditions

 Does ethylene polymerize under mild conditions in the absence of a catalyst (initiator)?

- ✓ No, in the absence of a catalyst, ethylene molecules would need to collide at very high energy in order to react with each other

 How does the molecular-level structure of these polymers influence their physical properties?

- ✓ The structure (e.g., extent of branching) determines how the individual polymer molecules can orient (or “pack”) in the solid state.
- ✓ polymers are biodegradable.

509. Which of the following statements is not true about low density polythene?

- A. Tough B. Hard C. Poor conductor of electricity D. Highly branched structure

Brief:

- ❖ Low density polythene: It is obtained by the polymerisation of ethane under high pressure of 1000 to 2000 atmosphere at a temperature of 350 K to 570 K in the presence of traces of dioxygen or a peroxide initiator (catalyst). They are flexible in nature.

510. Which of the following are characteristics of thermosetting polymers?

- A. Heavily branched cross linked polymers.
- B. Linear slightly branched long chain molecules.
- C. Become infusible on moulding so cannot be reused.
- D. Soften on heating and harden on cooling, can be reused.

Brief:

- ❖ Thermosetting polymers: These polymers are cross linked or heavily branched molecules, which on heating undergo extensive cross linking in moulds and again become infusible. These cannot be reused.

511. Which of the following polymers are thermoplastic?

- A. Teflon B. Natural rubber C. Neoprene D. Polystyrene

Brief: Thermoplastic polymers: These are the linear or slightly branched long chain molecules capable of repeatedly softening on heating and hardening on cooling. These polymers possess intermolecular forces of attraction intermediate between elastomers and fibres. Some common thermoplastics are polythene, polystyrene, polyvinyls, Teflon etc.

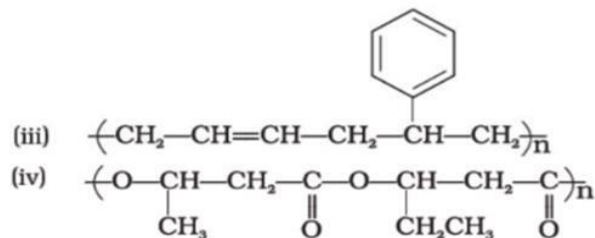
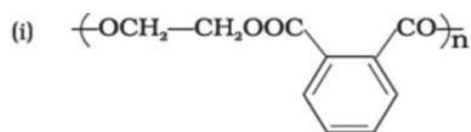
512. Vulcanization makes rubber ____.

- A. more inelastic B. soluble in inorganic solvent C. crystalline D. more stiff

Brief:

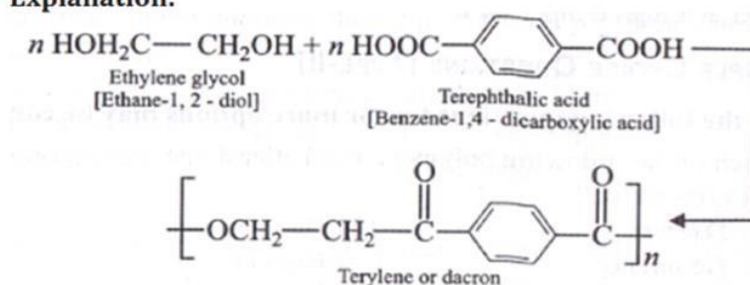
- ✓ To improve upon physical properties of natural rubber a process of vulcanization is carried out. This process consists of heating a mixture of raw rubber with sulphur and an appropriate additive at a temperature range between 373 K to 415 K. On vulcanization, sulphur forms cross links at the reactive sites of double bonds and thus the rubber gets stiffened.

513. In which of the following polymers ethylene glycol is one of the monomer units?



(i)

Explanation:



514. Match the polymer of column I with correct monomer of column II

Column I	Column II
(i) High density polythene	(a) Isoprene
(ii) Neoprene	(b) Tetrafluoroethene
(iii) Natural rubber	(c) Chloroprene
(iv) Teflon	(d) Acrylonitrile
(v) Acrilan	(e) Ethene

Explanation:

Column I	Column II
(i) High density polythene	It is formed when addition poly-merisation of ethene takes place in a hydrocarbon solvent in the presence of a catalyst
(ii) Neoprene	Neoprene or polychloroprene is formed by the free radical polymerisation of chloroprene
(iii) Natural rubber	Natural rubber may be considered as a linear polymer of isoprene (2-methyl-1, 3-butadiene)
(iv) Teflon	Teflon is manufactured by heating tetrafluoroethene with a free radical or persulphate catalyst at high pressures.
(v) Acrilan	The addition polymerisation of acrylonitrile in presence of a peroxide catalyst leads to the formation of polyacrylonitrile.

515. Match the polymer of column I with correct monomer of column II

Column I	Column II
(i) Nylon 6	(a) Polyvinyl chloride
(ii) PVC	(b) Polyacrylonitrile
(iii) Acrilan	(c) Polycaprolactum
(iv) Natural rubber	(d) Low density polythene
(V) LDP	(e) <i>cis</i> -polyisoprene

Explanation:

Column I	Column II
(i) Nylon 6	It is obtained by heating caprolactum with water at a high temperature.
(ii) PVC	Polymer of vinyl chloride
(iii) Acrilan	The addition polymerisation of acrylonitrile in presence of a peroxide catalyst leads to the formation of polyacrylonitrile.
(iv) Natural rubber	Natural rubber may be considered as a linear polymer of isoprene (2-methyl-1, 3-butadiene) and is also called as <i>cis</i> -1, 4 polyisoprene.
(v) LDP	The low-density polythene (LDP) obtained through the free radical addition and H-atom abstraction has highly branched structure.

Periodic Table of the Elements

1 IA H Hydrogen 1.008 1																		2 IIA He Helium 4.0026 2																		13 IIIA B Boron 10.81 2.5																		14 IVA C Carbon 12.011 2.5																		15 VA N Nitrogen 14.007 2.5																		16 VIA O Oxygen 15.999 2.4																		17 VIIA F Fluorine 18.998 2.7																		18 VIIIA Ne Neon 20.180 2.8																																																																																																																																																																																																					
3 Li Lithium 6.94 2.3																		4 Be Beryllium 9.012 2.3																		5 B Boron 10.81 2.5																		6 C Carbon 12.011 2.5																		7 N Nitrogen 14.007 2.5																		8 O Oxygen 15.999 2.4																		9 F Fluorine 18.998 2.7																		10 Ne Neon 20.180 2.8																																																																																																																																																																																																					
11 Na Sodium 22.98976928 2.8.1																		12 Mg Magnesium 24.305 2.8.2																		13 Al Aluminum 26.9815386 2.8.3																		14 Si Silicon 28.0855 2.8.4																		15 P Phosphorus 30.973762 2.8.5																		16 S Sulfur 32.06 2.8.6																		17 Cl Chlorine 35.45 2.8.7																		18 Ar Argon 39.948 2.8.8																																																																																																																																																																																																					
19 K Potassium 39.0983 2.8.9.1																		20 Ca Calcium 40.078 2.8.9.2																		21 Sc Scandium 44.955908 2.8.9.2																		22 Ti Titanium 47.88 2.8.9.2																		23 V Vanadium 50.9415 2.8.9.2																		24 Cr Chromium 51.9961 2.8.9.2																		25 Mn Manganese 54.938044 2.8.9.2																		26 Fe Iron 55.845 2.8.9.2																		27 Co Cobalt 58.9332 2.8.9.2																		28 Ni Nickel 58.6934 2.8.9.2																		29 Cu Copper 63.546 2.8.9.2																		30 Zn Zinc 65.38 2.8.9.2																		31 Ga Gallium 69.723 2.8.9.2																		32 Ge Germanium 72.630 2.8.9.2																		33 As Arsenic 74.9216 2.8.9.2																		34 Se Selenium 78.9718 2.8.9.2																		35 Br Bromine 79.904 2.8.9.2																		36 Kr Krypton 83.798 2.8.9.2																	
37 Rb Rubidium 85.4678 2.8.9.2																		38 Sr Strontium 87.62 2.8.9.2																		39 Y Yttrium 88.90584 2.8.9.2																		40 Zr Zirconium 91.224 2.8.9.2																		41 Nb Niobium 92.90638 2.8.9.2																		42 Mo Molybdenum 95.95 2.8.9.2																		43 Tc Technetium (98) 2.8.9.2																		44 Ru Ruthenium 101.07 2.8.9.2																		45 Rh Rhodium 102.91 2.8.9.2																		46 Pd Palladium 106.42 2.8.9.2																		47 Ag Silver 107.8682 2.8.9.2																		48 Cd Cadmium 112.411 2.8.9.2																		49 In Indium 114.818 2.8.9.2																		50 Sn Tin 118.710 2.8.9.2																		51 Sb Antimony 121.757 2.8.9.2																		52 Te Tellurium 127.60 2.8.9.2																		53 I Iodine 126.905 2.8.9.2																		54 Xe Xenon 131.29 2.8.9.2																	
55 Cs Cesium 132.90545196 2.8.9.2																		56 Ba Barium 137.327 2.8.9.2																		57-71 Lanthanides																		72 Hf Hafnium 178.49 2.8.9.2																		73 Ta Tantalum 180.94788 2.8.9.2																		74 W Tungsten 183.84 2.8.9.2																		75 Re Rhenium 186.207 2.8.9.2																		76 Os Osmium 190.23 2.8.9.2																		77 Ir Iridium 192.222 2.8.9.2																		78 Pt Platinum 195.084 2.8.9.2																		79 Au Gold 196.96657 2.8.9.2																		80 Hg Mercury 200.59 2.8.9.2																		81 Tl Thallium 204.38 2.8.9.2																		82 Pb Lead 207.2 2.8.9.2																		83 Bi Bismuth 208.9804 2.8.9.2																		84 Po Polonium (209) 2.8.9.2																		85 At Astatine (210) 2.8.9.2																		86 Rn Radon (222) 2.8.9.2																	
87 Fr Francium (223) 2.8.9.2																		88 Ra Radium (226) 2.8.9.2																		89-103 Actinides																		104 Rf Rutherfordium (261) 2.8.9.2																		105 Db Dubnium (268) 2.8.9.2																		106 Sg Seaborgium (269) 2.8.9.2																		107 Bh Bohrium (270) 2.8.9.2																		108 Hs Hassium (271) 2.8.9.2																		109 Mt Meitnerium (272) 2.8.9.2																		110 Ds Darmstadtium (281) 2.8.9.2																		111 Rg Roentgenium (282) 2.8.9.2																		112 Cn Copernicium (285) 2.8.9.2																		113 Nh Nihonium (286) 2.8.9.2																		114 Fl Flerovium (289) 2.8.9.2																		115 Mc Moscovium (290) 2.8.9.2																		116 Lv Livermorium (293) 2.8.9.2																		117 Ts Tennessine (294) 2.8.9.2																		118 Og Oganesson (294) 2.8.9.2																	
57 La Lanthanum 138.90547 2.8.9.2																		58 Ce Cerium 140.12 2.8.9.2																		59 Pr Praseodymium 140.90766 2.8.9.2																		60 Nd Neodymium 144.24 2.8.9.2																		61 Pm Promethium (145) 2.8.9.2																		62 Sm Samarium 150.36 2.8.9.2																		63 Eu Europium 151.964 2.8.9.2																		64 Gd Gadolinium 157.25 2.8.9.2																		65 Tb Terbium 158.92534 2.8.9.2																		66 Dy Dysprosium 162.50087 2.8.9.2																		67 Ho Holmium 164.93033 2.8.9.2																		68 Er Erbium 167.259 2.8.9.2																		69 Tm Thulium 168.93032 2.8.9.2																		70 Yb Ytterbium 173.05468 2.8.9.2																		71 Lu Lutetium 174.967 2.8.9.2																																																																							
89 Ac Actinium (227) 2.8.9.2																		90 Th Thorium 232.0377 2.8.9.2																		91 Pa Protactinium 231.036888 2.8.9.2																		92 U Uranium 238.02891 2.8.9.2																		93 Np Neptunium (237) 2.8.9.2																		94 Pu Plutonium (244) 2.8.9.2																		95 Am Americium (243) 2.8.9.2																		96 Cm Curium (247) 2.8.9.2																		97 Bk Berkelium (247) 2.8.9.2																		98 Cf Californium (251) 2.8.9.2																		99 Es Einsteinium (252) 2.8.9.2																		100 Fm Fermium (257) 2.8.9.2																		101 Md Mendelevium (258) 2.8.9.2																		102 No Nobelium (259) 2.8.9.2																		103 Lr Lawrencium (260) 2.8.9.2																																																																							