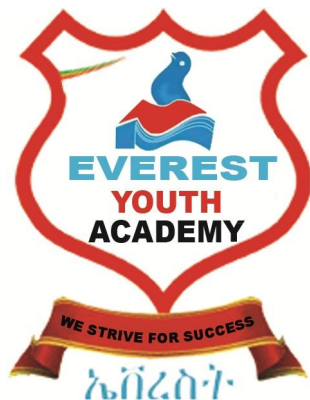


EVEREST YOUTH ACADEMY



SUPPLEMENTARY CHEMISTRY MODULE WITH DETAILED NOTES AND QUESTIONS

April/ 2015 E.C

Ethiopia

Contents

1	FUNDAMENTAL CONCEPTS IN CHEMISTRY FOR GRADE 11th	1
1.1	FUNDAMENTAL CONCEPTS IN CHEMISTRY	1
1.2	Questions.....	3
1.3	ANSWER.....	4
2	ATOMIC STRUCTURE AND PERIODIC TABLE FOR GRADE 11 th	5
2.1	ATOMIC STRUCTURE.....	5
2.1.1	Postulates of Dalton's Atomic Theory	5
2.1.2	Postulates of Modern Atomic Theory	5
2.1.3	Discovery of the Electron	6
2.1.4	Radioactivity and Discovery of Nucleus	6
2.1.5	Discovery of Nucleus	6
2.1.6	Rutherford's Conclusion:.....	7
2.1.7	Discovery of Neutron	7
2.2	MAKE UP OF THE NUCLEUS	7
2.2.1	Isotopes	7
2.3	ELECTROMAGNETIC RADIATION (EMR).....	8
2.3.1	The Quantum Theory and Photon	8
2.3.2	The Photoelectric Effect	8
2.3.3	Atomic Spectra.....	8
2.4	The Bohr Model of the Hydrogen Atom.....	9
2.4.1	Limitations of the Bohr Theory	9
2.5	THE QUANTUM MECHANICAL MODEL OF THE ATOM	9
2.5.1	The Heisenberg's Principle	9
2.5.2	Quantum Numbers.....	9
2.6	ELECTRONIC CONFIGURATIONS	10
2.7	ELECTRONIC CONFIGURATIONS AND THE PERIODIC TABLE OF THE ELEMENTS	11
2.7.1	The Modern Periodic Table	11
2.7.2	Classification of the Elements	11

2.7.3	Periodic Properties	11
2.7.4	Advantages of Periodic Classification of the Elements	12
2.8	Questions.....	13
2.9	ANSWER.....	16
3	CHEMICAL BONDING AND HYBERDIZATION FOR GRADE 11 th	17
3.1	Chemical bond.....	17
3.2	Ionic Bond.....	17
3.3	Covalent compound that violet the Octate rule.....	18
3.3.1	Coordinate covalent bond /dative/donor-acceptor	19
3.3.2	Polar and non-polar covalent bonds	19
3.3.3	Difference between ionic compound and covalent compound	19
3.4	Types of electron pair arrangement.....	20
3.5	Intermolecular forces in covalent compound.....	20
3.5.1	Difference between intermolecular force and intramolecular forces	21
3.6	Chemical Bonding theories	21
3.6.1	Valance bond theories (VBT)	21
3.6.2	Hybridization of orbitals.....	22
3.6.3	Molecular orbital theory	22
3.7	Types of solids	23
3.7.1	Types of crystals	24
3.8	Questions.....	25
3.9	Answer	29
4	UNIT FOUR CHEMICAL KINETICS FOR GRADE 11 th	30
4.1	Chemical Kinetics	30
4.2	Reaction Rates.....	30
4.3	Factors Affecting Reaction Rate Constants	31
4.4	Concentration and Rate.....	31
4.5	Order of Reactions.....	32
4.6	Reaction Mechanisms	34
4.7	Rate Laws and Rate Determining Steps	35

4.8	Activation Energy.....	35
4.9	EXERCISE.....	36
4.10	ANSWERS	41
5	UNIT FIVE CHEMICAL EQUILIBRIUM AND PHASE EQUILIBRIUM FOR GRADE	
11 th	41	
5.1	CHEMICAL EQUILIBRIUM	42
5.2	Law of Mass Action	43
5.3	Factors affecting chemical equilibrium	46
5.4	Question	48
5.5	ANSWER.....	51
6	UNIT SIX HYDRO CARBONS AND DERIVATIVE HYDRO CARBON FOR GRADE	
10 TH & 11 TH		52
6.1	HYDRO CARBONS AND DERIVATIVE HYDRO CARBON	53
6.2	IUPAC name organic compound.....	53
6.3	Chemical properties of aliphatic hydrocarbons	54
6.4	Aromatic Hydrocarbons.....	55
6.5	Natural source of Hydrocarbons	55
6.6	Alcohol, Aldehyde, Ketone Carboxylic acid & Ester	56
6.7	Question	58
6.8	ANSWER.....	61
7	UNIT ONE SOLUTIONS FOR GRADE 12 th	62
7.1	INTRODUCTION	62
7.2	Suspension and Colloids.....	62
7.3	Types of Solutions.....	62
7.4	The Solution Process	63
7.4.1	Liquid Solutions and Inter-particle Forces of Attractions.....	63
7.4.2	The Rate of Dissolution	64
7.4.3	Energy Changes in the Solution Process	64
7.4.4	Hydration of Ionic Solids in Water	64
7.4.5	Solubility as an Equilibrium Process.....	64

7.4.6	Ways of Expressing Concentrations of Solution	65
7.4.7	Mass Percentage, ppm, and ppb of Solute	65
7.4.8	Mole Fraction (X)	65
7.4.9	Molarity (M)	65
7.4.10	Normality (N).....	66
7.4.11	Molality (m).....	66
7.4.12	Diluting Solutions	66
7.5	DESCRIBING REACTIONS IN SOLUTION	66
7.5.1	Molecular Equation	66
7.5.2	Ionic Equation	66
7.6	COLLIGATIVE PROPERTIES OF SOLUTION	67
7.6.1	Vapor Pressure Lowering	67
7.6.2	Boiling Point Elevation	67
7.6.3	Freezing Point Depression	68
7.6.4	Osmosis and Osmotic Pressure	68
7.6.5	Colligative Properties of Electrolytic Solutions	68
7.7	Questions.....	70
7.8	ANSWERS.....	71
8	UNIT TWO ACID – BASE EQUILIBIRA FOR GRADE 12 th	73
8.1	INTRODUCTION	73
8.2	Acid - Base Equilibrium	73
8.2.1	Brønsted-Lowry Concept of Acids and Bases	74
8.2.2	Auto-ionization of Substances	74
8.3	Ionization of Water.....	75
8.4	The pH scales	75
8.5	Measures of the Strengths of Acids and Bases in Aqueous Solution.....	76
8.5.1	Acid Dissociation Constant, K_a	76
8.5.2	Base dissociation constant, K_b	76
8.6	The Common ion Effect.....	76
8.7	Buffer Solutions	76

8.8	HYDROLYSIS OF SALTS	77
8.9	Acid–Base Indicators.....	77
8.10	Equivalents of Acids and Bases.....	77
8.11	Acid–Base Titrations.....	77
8.11.1	Strong Acid-Strong Base Titrations	78
8.11.2	Weak Acid-Strong Base Titrations	78
8.12	Questions	79
8.13	ANSWERS	80
9	UNIT 3 INTRODUCTION TO THERMODYNAMICS	82
9.1	INTRODUCTION TO THERMODYNAMICS	82
9.2	THE FIRST LAW OF THERMODYNAMICS AND SOME THERMODYNAMIC QUANTITIES	83
	Internal Energy (E)	83
9.2.1	Heat (Q)	84
9.2.2	Work (W)	84
	The First Law of Thermodynamics	85
	Thermochemistry	85
9.2.3	Heats of Reactions	85
9.2.4	Standard States	86
	Hess’s Law	86
9.2.5	Entropy and Spontaneous Process.....	86
9.2.6	The Second Law of Thermodynamics	87
9.3	Question	88
9.4	Answer	90
10	UNIT 4 ELECTRO CHEMISTRY FOR GRADE 12 th	91
10.1	ELECTROCHEMISTRY	91
10.2	Electrochemical cells	91
10.3	Types of electrochemical cells.....	91
10.3.1	Electrolytic cell	92
10.3.2	Galvanic cell	92

10.4	Electrode potential	92
10.5	Standard electrode potential (E°).....	92
10.6	Electrochemical series.....	93
10.7	Cell potential or emf of a cell	93
10.8	Cell Diagram or Representation of a Cell	93
10.9	4.8. Salt bridge	94
10.10	Spontaneity of a reaction	94
10.11	NERNST EQUATION	94
10.12	Concentration cells	95
10.13	BATTERIES.....	96
10.14	CORROSION	97
10.15	Questions.....	98
10.16	Answers.....	102
11	UNIT 5 CHEMISTRY OF SOME ELEMENTS AND COMPOUNDS FOR GRADE 10 th &12 th	103
11.1	Chemistry of Some Elements and Compounds	103
11.2	Metallurgy	103
11.3	Important metals	103
11.4	Alloys	104
11.5	Important compounds.....	105
11.6	Questions Grade 10 th &12 th (important inorganic metals and compounds)	105
11.7	Answer to review question Grade 10 th &12 th on important metals and compounds 109	
12	UNIT 6 POLYMER CHEMISTRY FOR GRADE 12 th	111
12.1	6. POLYMER	111
12.2	Question.....	113
12.3	Answers for unit 6 polymer	114
13	Grade 10 SOME IMPORTANT IN ORGANIC COMPOUNDS	114
13.1	SOME IMPORTANT IN ORGANIC COMPOUNDS.....	115
13.2	Question.....	120

13.3	Grade 10 SOME IMPORTANT IN ORGANIC COMPOUNDS.....	122
14	GRADE 9 THE PHYSICAL STATES OF MATTER.....	123
14.1	The Kinetic Theory of Matter.....	123
14.2	THE GASEOUS STATE	123
14.3	The Gas Laws	123
14.4	Energy Changes in Liquids.....	124
14.5	Question.....	125
14.6	ANSWER GRADE 9 THE PHYSICAL STATES OF MATTER	126

1 FUNDAMENTAL CONCEPTS IN CHEMISTRY FOR GRADE 11th

Unit Outcomes

At the end of this unit, you should be able to:

- ✓ understand the scope of chemistry;
- ✓ select and use appropriate SI units;
- ✓ understand the causes of uncertainty in measurement;
- ✓ express the result of any calculation involving experimental data to the appropriate number of decimal places or significant figures;
- ✓ use scientific methods in solving problems;

1.1 FUNDAMENTAL CONCEPTS IN CHEMISTRY

Chemistry is the science that deals with matter and the changes that it undergoes. It is a study of the composition, structure, and properties of matter, and of the changes that occur in matter.

There are two kinds of physical properties, namely, extensive, and intensive physical properties.

- ✓ Extensive physical properties are the properties, which depend on the amount or quantity of sample and therefore, can vary from sample to sample. The extensive property of a piece of copper wire, for instance, includes its length, diameter, mass, and electrical resistance.
- ✓ Intensive physical properties are properties which do not depend on the amount of a substance present. The intensive properties of a piece of copper wire include its density, color, melting point, and hardness. Intensive properties are useful in distinguishing between different substances because they do not vary from sample to sample.

There are several branches of chemistry, the major branches are, Inorganic chemistry, Organic chemistry, Physical chemistry, and Analytical chemistry.

Basic Property

Length

Mass

Time

Temperature

Electric Current

Amount of substance

Luminous intensity

Metric Unit

meter (m)

kilogram (kg)

second (s)

kelvin (K)

ampere (A)

mole (mol)

candela (cd)

Prefix Name	Affect on unit
yotta (y)	$\text{unit} \times 10^{24}$
zetta (Z)	$\text{unit} \times 10^{21}$
exa (E)	$\text{unit} \times 10^{18}$
peta (P)	$\text{unit} \times 10^{15}$
tera (T)	$\text{unit} \times 10^{12}$ (1,000,000,000,000)
giga (G)	$\text{unit} \times 10^9$ (1,000,000,000)
mega (M)	$\text{unit} \times 10^6$ (1,000,000)
kilo (k)	$\text{unit} \times 10^3$ (1,000)
hecto (h)	$\text{unit} \times 10^2$ (100)
deca (deca) (da)	$\text{unit} \times 10$
deci (d)	$\text{unit} \div 10$
centi (c)	$\text{unit} \div 100$
milli (m)	$\text{unit} \div 10^3$ (1,000)
micro (μ)	$\text{unit} \div 10^6$ (1,000,000)
nano (n)	$\text{unit} \div 10^9$ (1,000,000,000)
pico (p)	$\text{unit} \div 10^{12}$ (1,000,000,000,000)
femto (f)	$\text{unit} \div 10^{15}$
atto (a)	$\text{unit} \div 10^{18}$
zepto (z)	$\text{unit} \div 10^{21}$
yocto (y)	$\text{unit} \div 10^{24}$

Precision → agreement or closeness among several measurement of the same quantity.

Accuracy → closely individual measurements agree with the correct value.

A point or dot (•) used to separate the whole number part from the fractional part of a number is called a decimal point.

To express a value to the n th decimal place, look at the values of the $(n + 1)^{\text{th}}$ digit, as stated in the following rules.

1. If the $(n + 1)^{\text{th}}$ digit is 4 or less, leave the n^{th} digit unchanged. For instance, 369.648 rounds to 369.6 if we want one decimal place.
2. If the $(n + 1)^{\text{th}}$ digit is greater than 5 round up the n^{th} digit. For instance, 369.758 rounds to 369.76 if we want two decimal places.
3. If the digit to be removed is 5, the preceding number increases by one unit if it is odd and remains unchanged if it is even. For example, 17.75 rounds to 17.8, but 17.65 rounds to 17.6. Note that if 5 is followed only by zeros, the left-most digit is unchanged. However, if the 5 is followed by non-zeros, the final digit is increased by 1. For example, 17.6500 rounds to 17.6, but 17.6513 rounds to 17.7.

- ✓ If the zeros precede the first non-zero digit, they are not significant. Such zeros merely locate the decimal point.
- ✓ If the zeros are between non-zero digits, they are significant.
- ✓ If the zeros follow non-zero digits, there is ambiguity if no decimal point is given.
- ✓ Non-zero digits are always significant.
- ✓ For addition & subtraction for significant figures, use least number of decimal places.
- ✓ For multiplication & division for significant figures, use least number of significant digits.
- ✓ The scientific method is a general approach to problems. It involves making observations, collecting information/data, seeking patterns in the observations, formulating hypothesis to explain the observations, and testing these hypothesis by further experiments. If a hypothesis successfully possesses many tests, it becomes a theory, which is a tested explanation of hypothesis.
- ✓ 1. Reaction vessels, example Beakers, flasks, boiling tubes and test tubes.
2. Measuring equipment's, example Pipettes, burettes, balances, and thermometers.
3. Support and heating devices, example Stand and clamp, tripod and gauze, spirit burner and Bunsen burner.
- ✓ Each report must include: title, objective, materials and (equipment) used, procedure, observation, result, discussion, and conclusion

1.2 Questions

- 1) The number $2.03 \times 10^{-2} \text{m} + 4.26 \text{cm} + 0.16 \text{dm}$ has _____ significant figures.
A. 1 B. 3 C. 2 D. 4
- 2) Of the following are causes of uncertainty except
A. the measuring device
B. the environment where measurement is being made
C. the person doing the measurement
D. the number of measurements
- 3) which of the following branch of chemistry involves quantitative determination & qualitative identification
A. Organic chemistry C. Analytical chemistry
B. inorganic chemistry D. Physical chemistry
- 4) Which one of the following is the smallest number of atoms?
A. 1g of H_2 B. 1g F_2 C. 1g O_2 D. 1g N_2
- 5) A fraction of 10^{-12} represents a prefix
A. micro B. Nano C. Pico D. milli
- 6) Which of the following the best step in scientific methods?
A. Hypothesis $\rightarrow$ law $\rightarrow$ observation $\rightarrow$ experiment $\rightarrow$ theory

- B. law → theory → experiment → Hypothesis → observation
 C. observation → Hypothesis → experiment → theory → law
 D. experiment → Hypothesis → observation → theory → law
- 7) Which of the following measurement of length is the shortest?
 A. 1584cm B. 0.0128km C. 17926mm D. 14m
- 8) Which of the following is an exact number?
 A. 10.25g B. 4.000Kg C. 7chairs D. 60seconds
- 9) A chemistry teacher has given a sample of potassium chloride to four students in school laboratory. W, X, Y&Z. Each student is asked to weigh the sample and record his or her results. The accepted value is 12.72g. Their results for three trials given below

Trials	Student W	Student X	Student Y	Student Z
1	12.50g	12.50g	12.72g	12.41g
2	12.48g	12.77g	12.74g	12.72g
3	12.51g	12.83g	12.70g	12.55g

- Which set of data is the most accurate measurement?
 A. Student Z B. Student W C. Student X D. Student Y
- 10) When numbers are added, the limiting term is the one with the ____ number of ____
 A. largest, decimal places C. largest, significant digits
 B. smallest, decimal l places D. smallest, significant digits
- 11) The scientific notation of 0.000840 is _____.
 A. 8.40×10^{-3} B. 8.40×10^4 C. 8.40×10^{-4} D. 8.4×10^4

1.3 ANSWER

1. C	2. D	3. C	4. B	5. C	6. C
7. B	8. D	9. C	10. B	11. C	

2 ATOMIC STRUCTURE AND PERIODIC TABLE FOR GRADE 11th

Unit Outcomes

At the end of this unit, you should be able to:

- ✓ understand the historical development of atomic structure;
- ✓ explain the experimental observations and inferences made by some famous scientists to characterize the atom;
- ✓ list and describe the subatomic particles;
- ✓ explain the terms “atomic mass” and “isotope;”
- ✓ understand electromagnetic radiation, atomic spectra, and Bohr’s models of the atom;
- ✓ do calculations involving atomic structure;
- ✓ describe the quantum mechanical model of the atom, related postulates and principles;
- ✓ demonstrate an understanding of periodic law and how electronic configuration of atoms are related to the orbital diagrams and can explain periodic trends

2.1 ATOMIC STRUCTURE

Democritus

- discontinuous theory of matter
- Believed that all matter was composed of very tiny, indivisible particles.

Aristotle

- continuous theory of matter
- It would continue to divide infinitely.

2.1.1 Postulates of Dalton’s Atomic Theory

1. All substances are made up of very small particles called atoms
2. Atoms are indivisible and indestructible
3. All atoms of the same element are exactly alike in size, mass, and other properties.
4. Atoms of any one element are quite different from atoms of all other elements in their properties
5. Atoms of different elements combine in a fixed small whole number ratio form compounds.
6. In a given compound, the relative types and numbers of atoms are constant.

2.1.2 Postulates of Modern Atomic Theory

1. Atoms are divisible that can be sub divided in to electrons, protons and neutrons.
2. Atoms of the same element have identical chemical properties.
3. Atoms of the same element may not be exactly alike.
4. Atoms of different elements have different chemical properties

5. Atoms of two or more elements interact chemically the atoms unite in certain simple ratios to form compounds
6. In certain cases, the atoms of an element unite to form small independently existing units called molecules.

2.1.3 Discovery of the Electron

J.J. Thomson established clearly that cathode rays consist of negatively-charged particles. These negatively charged particles were constituents of every kind of atom. We now call these particles electrons.

Thomson was able to calculate the ratio of the mass of an electron to its charge.

$$e^-/m = 1.76 \times 10^8 \text{ C/g}$$

Robert A. Millikan determined the charge on the electron by observing the behavior of electrically charged oil drops in an electric field. $e^- = 1.602 \times 10^{-19} \text{ C}$

$$\text{Mass of the electron} = \frac{1.602 \times 10^{-19} \text{ C}}{1.76 \times 10^8 \text{ C/g}} = 9.11 \times 10^{-28} \text{ g} = 9.11 \times 10^{-31} \text{ kg}$$

An electron is a fundamental particle of an atom carrying a negative charge and having a very small mass.

2.1.4 Radioactivity and Discovery of Nucleus

Radioactivity is the spontaneous emission of radiation from the unstable nuclei of certain isotopes. For example, $^{238}_{92}\text{U}$

Three types of rays were emission from radioactive substances.

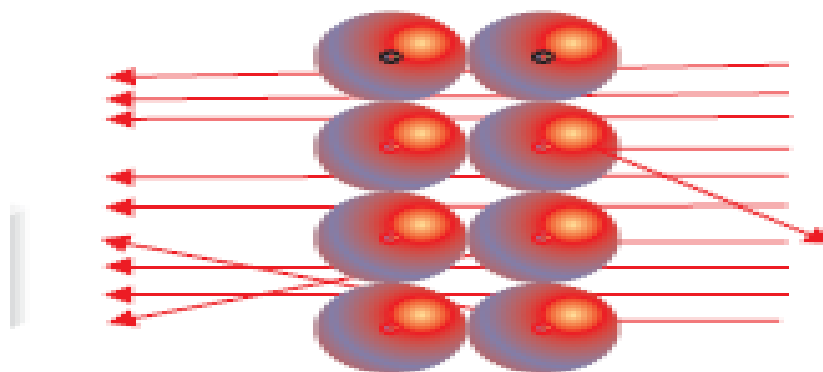
Alpha (α) particles consist of positively charged particles. They are deflected by positively charged plates. E.g. He^{2+} .

Beta (β) particles are electrons of nuclear origin that are deflected by negatively charged plates.

Gamma (γ) rays have no charge and are not affected by an external electric or magnetic field. Consist of high-energy rays like X-rays,

2.1.5 Discovery of Nucleus

Ernest Rutherford carried out a series of experiments using very thin foils of gold and other metals as targets for α -particles from a radioactive source



E.g. **Rutherford's α – particles experiment.**

2.1.6 Rutherford's Conclusion:

1. Since most of the α -particles passed through the gold foil undeflected due to most of the space in an atom is empty.
2. Some of the α -particles were deflected by small angles due to the presence of a heavy positive center in the nucleus of the atom.
3. A very few α -particles were deflected or bounced back to the source due to the mass of an atom is concentrated in the nucleus.

Rutherford proposed that protons constitute the positively charged matter in the nuclei of all atoms

2.1.7 Discovery of Neutron

James Chadwick bombarded a thin foil of beryllium with α -particles of a radioactive substance. He then observed that highly penetrating rays, consisting of electrically neutral particles of a mass approximately equal to that of the proton, were produced. These neutral particles are called neutrons.

A neutron is a subatomic particle carrying no charge and having a mass of 1.675×10^{-24} g. This mass is almost equal to that of a proton or of a hydrogen atom.

2.2 MAKE UP OF THE NUCLEUS

The nucleus of an atom contains protons and neutrons.

2.2.1 Isotopes

- ✓ Are the atoms of the same element that have the same atomic number but different masses number.
- ✓ Have the same number of protons and electrons but different number of neutrons.

- ✓ Have the same chemical properties.

2.3 ELECTROMAGNETIC RADIATION (EMR)

James Clerk Maxwell proposed that light consists of electromagnetic waves.

Electromagnetic radiation is the emission and transmission of energy in the form of electromagnetic waves.

The wave properties of electromagnetic radiation are described by two interdependent variables, frequency and wavelength.

$$C = \lambda \times \nu \text{ in vacuum, } C = 3.00 \times 10^8 \text{ m s}^{-1}$$

2.3.1 The Quantum Theory and Photon

The central idea of the quantum theory is that energy is not continuous but comes in discrete packets the name quantum is given to such a packet of energy.

Max Planck proposed that energy can be gained or lost only in whole number multiples of the quantity $h\nu$.

$$E = nh\nu \text{ where } h \text{ is Planck's constant} = 6.62 \times 10^{-34} \text{ J.s}$$

2.3.2 The Photoelectric Effect

Is light shining on a clean metallic surface can cause the surface to emit electrons?

For each metal, there is a minimum frequency of light (threshold frequency) below which no electrons are emitted, regardless of how intense the beam of light.

$$\Delta E = h\nu = h\nu_0 + \frac{1}{2}mv^2$$

Where ν_0 = the threshold frequency,

m_e = the mass of the electron and

V = velocity of the emitted electron.

ΔE = energy of a quantum of radiation

$h\nu$ = that goes into ejecting a photoelectron out of the metal

$h\nu_0$ = work function

$\frac{1}{2}mv^2$ = the kinetic energy

2.3.3 Atomic Spectra

- Continuous spectrum produced** by sun light because one color merges in to the next without any gaps or missing colors. e.g. sun light
- Line spectrum** excited atoms contain excess energy, which they release by emitting light of various wavelengths to produce.

2.4 The Bohr Model of the Hydrogen Atom

Bohr proposed the following three postulates for his model.

1. The hydrogen atom has only certain allowable energy levels, called stationary states.
2. The atom does not radiate energy while in one of its stationary states.
3. The electron moves to another stationary state (*orbit*) only by absorbing or emitting a photon whose energy equals the difference in the energy between the two states.

$$E_{\text{ph}} = E_f - E_i = h\nu$$

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

A very useful result from Bohr's work is an equation for calculating the energy levels of an atom,

$$E_n = -2.18 \times 10^{-18} \text{ J} \times \frac{1}{n^2}$$

2.4.1 Limitations of the Bohr Theory

The Bohr model was adequate for explaining atoms and ions with one electron such as H, He⁺ and Li²⁺. However, the Bohr model failed to predict the wavelengths of spectral lines of atoms more complicated than hydrogen.

2.5 THE QUANTUM MECHANICAL MODEL OF THE ATOM

Schrödinger equation leads to a series of solutions that describe the allowed energy states of the electron. A wave function provides information about an electron location in space when it is in a certain allowed energy state.

2.5.1 The Heisenberg's Principle

States that it is impossible to know simultaneously both the momentum and the position of a particle with certainty.

$$\Delta x \times \Delta p \leq h/4\pi$$

2.5.2 Quantum Numbers

The four quantum numbers are:

1. The principal quantum number (*n*)

It gives the following information:

- i. Relative size of the orbital

Size of orbital increases with the increase of principal quantum number *n*.

- ii. Energy of the orbital. Higher the *n* value, greater is the energy.

iii. Maximum number of electrons present in any shell given by the formula $2n^2$.

2. The azimuthal quantum number (l) /angular momentum or subsidiary quantum number/

- It is an integer having values from 0 to $(n - 1)$.

It gives the following information:

- Number of subshell present within any shell.
- It describes the shape of the orbital

3. The magnetic quantum number (ml)

- It is an integer having values from $-l$ to $+l$.
- *The number of possible ml values is $(2l + 1)$.*
- *It describes the orientation of the orbital*

4. The electron spins quantum number (ms)

- Has only two possible values, $+\frac{1}{2}$, $-\frac{1}{2}$ ($\uparrow\downarrow$).
- Suggests that electrons have a spinning motion.

Each sublevel is designated by a letter:

$l = 0$, is an s sublevel; $l = 1$, is a p sublevel; $l = 2$, is a d sublevel; $l = 3$, is a f sublevel
Number of orbitals = n^2

2.6 ELECTRONIC CONFIGURATIONS

The electron configuration for any atom follows the following three principles:

- ✓ **Aufbau Principle.** Electrons occupy the lowest-energy orbital available before entering the higher energy orbital.
- ✓ **Hund's Principle.** Equal energy orbitals (degenerate orbitals) are each occupied by a single electron before the second electrons of opposite spin enters the orbital. In other words, each of the three $2p$ orbitals ($2px$, $2py$ and $2pz$) will hold a single electron before any of them receives a second electron.
- ✓ **Pauli's Exclusion Principle.** No two electrons can have the same four quantum numbers. *i.e.*, they must differ in at least one of the four quantum numbers.
- ✓ **Ground State Electronic Configuration of the Elements** Two general methods are used to denote electron configurations. The subshell (sublevel) the letters s , p , d and f and principal quantum number.

Expected

Cr ($Z = 24$) = $[\text{Ar}] 4s^2 3d^4$

Cu ($Z = 29$) = $[\text{Ar}] 4s^2 3d^9$

Observed

$[\text{Ar}] 4s^1 3d^5$

$[\text{Ar}] 4s^1 3d^{10}$

2.7 ELECTRONIC CONFIGURATIONS AND THE PERIODIC TABLE OF THE ELEMENTS

2.7.1 The Modern Periodic Table

States that certain sets of physical and chemical properties recur at regular intervals (periodically) when the elements are arranged according to increasing atomic number.

2.7.2 Classification of the Elements

Elements are placed in the periodic table in accordance to valence electron entering the orbital of lowest energy. There are 18 groups and 7 periods in the modern periodic table.

A. Representative or main group elements:

These consist of all *s*- and *p*-block elements. The chemical properties of the representative elements are determined by the number of valence electrons in their atoms.

B. Transition elements:/*d*-block elements/

There are four series of transitional elements, *3d*, *4d*, *5d* and *6d* depending on the energy levels of *d*-orbitals.

C. Inner transition elements: / *f*-block elements/

There are two series of *f*-block elements, *4f* and *5f* series called lanthanides and actinides, respectively.

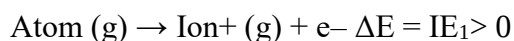
2.7.3 Periodic Properties

In this section you will examine some periodic atomic properties like atomic radii, ionization energies, electron affinities, electronegativity, and metallic character.

Atomic Size (Atomic Radii) is the distance between the nuclei of two adjacent atoms.

- ✓ Atomic radii of the A-group elements (main group element) tend to decrease from left to right in a period of the periodic table but increases down a group.

Ionization Energy (IE) is the amount of energy required to remove the outermost electron in an isolated gaseous atom or ion.



- ✓ The first ionization energy (IE_1) is the energy needed to remove an electron from the highest occupied sublevel of the gaseous atom.
- ✓ The second ionization energy (IE_2) removes the second electron.
- ✓ IE_2 is always larger than IE_1

Factors affecting ionization energy

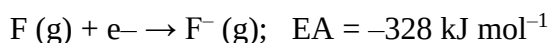
- ✓ Atomic radius
- ✓ Effective nuclear charge
- ✓ Screening effect
- ✓ Type of sub levels

✓ Electron Configuration

In general, ionization energy increases across a period; it is easier to remove an electron from an alkali metal than from a noble gas.

Electron Affinity (EA)

Is the energy released that occurs when an electron is added to a gaseous atom or ion for the formation of a gaseous negative ion?



- ✓ Smaller an atom, more negative is its electron affinity.
- ✓ The smaller the atom the closer an added electron can approach the atomic nucleus and the more strongly it is attracted to the nucleus.

Electronegativity

Is the ability of an atom to attract electrons towards to itself.

- ✓ Atoms of the elements in the upper right of the periodic table (small, non-metal atoms) attract bonding electrons most strongly. Therefore, they have the greatest electro negativities.
- ✓ Atoms of the elements to the left side of the table (large, metal atoms) have a weaker hold on electrons.
- ✓ the most electronegative elements are fluorine, oxygen and nitrogen
- ✓ Within a period of the periodic table, elements become more electronegative from left to right.
- ✓ Within a group, electronegativity decreases from top to bottom.

Metallic Character

- ✓ Is the elements ability to lose electrons to form cation?
- ✓ As one moves across a period from left to right the metallic character decreases.
- ✓ Down a group, the metallic character increases, due to the lesser attraction from the nucleus to the valence electrons.

2.7.4 Advantages of Periodic Classification of the Elements

Some of the advantages of periodic classification of elements are:

1. To predict the physical and chemical properties of the elements
2. To predict the physical and chemical properties of compounds
3. To predict the formula of compounds
4. The lanthanides and actinides are placed separately at the bottom of the periodic table.
5. The table is simple, systematic and easy way for remembering the properties of various elements

2.8 Questions

- Of the following assumptions of Dalton's atomic theory, the only one that remains essentially correct is
 - All atoms of an element are identical in mass
 - Atoms are indivisible and indestructible
 - Atoms are divisible and destructible
 - Atoms of an element combines in the ratios of small whole numbers to form compound
- The scattering of alpha particles by thin metal foils established that
 - The mass and positive charge of an atom is concentrated in a nuclei's
 - Electron is fundamental particles of all matter
 - All electrons have the same charge
 - Atoms is electrically neutral
- Which of the following element has the highest first ionization energy?
 - Carbon
 - Nitrogen
 - Oxygen
 - Fluorine
- Which one has the largest size?
 - Al^{3+}
 - Na^+
 - F^-
 - N^{3-}
- Elements in the same periodic group are characterized by the same
 - Electron affinity
 - Ionization energy
 - Number of shells
 - Atomic radius
- The most electronegativity elements are found
 - Left bottom of the periodic table
 - On the right bottom of the periodic table
 - At center of the periodic table
 - On the right top of the periodic table
- The power of an atom to attract electrons to itself when bonded to another atom is
 - Ionization energy
 - Electronegativity
 - Electron affinity
 - Electro positivity
- The minimum amount of energy required to remove an electron from the surface of a substance is.
 - Photoelectric effect
 - Work function
 - Electromagnetic radiation
 - Atomic spectra
- For an electron that has quantum number $n=4$ and $m_l=0$ which of the following is true.
 - It must have quantum number $n=0$
 - It must have the quantum number $l=0$

- C. It must have the quantum number $m_s = -1/2$
 D. It may have the quantum number $l = 0, 1, 2, 3$
10. What are the possible values of the azimuthal quantum number for an electron in the 4f sub shell?
- A. 2 B. 5 C. 3 D. 4
11. Which of the following is true about properties of cathode rays?
- A. They cast shadows
 B. They deflected by electric & magnetic field
 C. The heat a metal foil placed b/n the electrode
 D. They travel in straight lines
12. What did J.J Thomson experiment show?
- A. The behavior of cathode rays
 B. Cathode rays consisting negatively charged particles
 C. Cathode rays are beams of electrons
 D. all are correct
13. What is the difference b/n Br-79 & Br-81?
- A. Br-79 has two more protons than Br-81.
 B. Br-79 has two more neutrons than Br-81.
 C. Br-79 has two more electrons than Br-81.
 D. Br-79 has two more electrons & one more neutron than Br-81.
14. A Fe^{2+} ion has
- A. 26 protons, 28 neutrons & 24 electrons
 B. 24 protons, 28 neutrons & 26 electrons
 C. 26 protons, 30 neutrons & 26 electrons
 D. 26 protons, 26 electrons & 24 electrons]
15. Which of the following EMR has the highest energy?
- A. X-ray B. UV rays C. IR D. RW
16. The following are representative wave lengths in the IR, UV, and X-rays in EM spectrum with wave length $1.0 \times 10^{-6}\text{m}$, $1.0 \times 10^{-8}\text{m}$ & $1.0 \times 10^{-10}\text{m}$ respectively. The highest energy of a photon of each radiation is: -
- A. IR B. UV C. X-ray D. All are the same energy

17. The color in the visible region of the EM spectrum arranged increasing order of wave length. Which abbreviation written more easily?

A. VIBYGOR B. VIBGYOR C. YIBGVOR D. VOBGYIR

18. Which of the following electron transitions requires the highest energy to be absorbed by the hydrogen atom?

A. From $n=5$ to $n=4$ C. From $n=2$ to $n=3$
B. From $n=4$ to $n=5$ D. From $n=1$ to $n=2$

19. A neutral atom X at its ground state has half-filled 4d orbital. The atomic number of X is: -

A. 24 B. 59 C. 42 D. 74

20. Which set of quantum number (n , l , m_l , and m_s) is not permitted by the rules of quantum mechanics?

A. $n=2, l=0, m_l=0, m_s=+1/2$ C. $n=4, l=3, m_l=-3, m_s=+1/2$
B. $n=3, l=3, m_l=-1, m_s=-1/2$ D. $n=1, l=0, m_l=0, m_s=-1/2$

21. Of the following is not true about Electron configuration.

A. Building up principles- orbitals with the lowest energy filled first.
B. Pauli's exclusion principles – degenerate orbital's
C. Orbital diagram – determine the number of unpaired electrons in an atom.
D. Pauli's exclusion principles – limit the number of electrons in an orbital to two.

22. Which of the following is invalid electron configuration?

A. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ C. $1s^2 2s^2 2p^6 3s^2 3p^4$
B. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ D. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$

23. The sub shell that is found in every element is: -

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

24. The main advantage of periodic classification of an element is:-

A. Useful to predicting the formulas of the compounds
B. Useful for predicting the physical and chemical properties of elements.
C. Useful for predicting the behavior of many compounds.
D. all

25. Which atomic model in an atom, electrons follow a circular orbital around the nucleus?

A. Daltons model C. Bohr's model
B. Thomson's model D. Rutherford's model

26. When the atoms Na, Mg, Al, and K are arranged in order of increasing atomic radius, which is the correct order?

A. Na, Mg, Al, K B. Al, Mg, Na, K C. Na, K, Al, Mg D. K, Na, Mg, Al

2.9 ANSWER

1. D	2. A	3. B	4. D	5. C	6. D	7. B	8. B
9. D	10. C	11. D	12. D	13. B	14. A	15. A	16. C
17. B	18. D	19. C	20. B	21. B	22. D	23. A	24. D
25. C	26. B						

3 CHEMICAL BONDING AND HYBERDIZATION FOR GRADE 11th

Unit Outcomes

At the end of this unit, you should be able to:

- ✓ understand that a chemical bond is an attractive force between particles;
- ✓ demonstrate an understanding of the formation and general properties of substances containing ionic, covalent, and metallic bonds;
- ✓ draw Lewis structures for simple ionic and covalent compounds;
- ✓ understand the origin of polarity within molecules;
- ✓ describe the formation and nature of hydrogen bonds, dipole-dipole forces, and London forces;
- ✓ know the three different but related bonding models (Lewis model, Valence bond model and Molecular orbital model) and recognize the usefulness of the bonding theories in explaining and predicting molecular properties (bond angle, bond length, bond energy, etc.);
- ✓ explain how the properties of a substance (solid or liquid) depends on the nature of the particles present and the type of intermolecular forces;
- ✓ appreciate the importance of intermolecular forces in plant and animal life;
- ✓ explain how the Valence Shell Electron Pair Repulsion (VSEPR) model can be used to predict molecular shape;
- ✓ Know the types of crystalline solid (ionic, molecular, covalent network, or metallic) formed by a substance and describe their properties

3.1 Chemical bond

- ✓ holds atoms together

No	Types of chemical bond	Formation of bond	example
1	Ionic bond	Transfer of electrons	KCl, MgO, CaCl ₂ , etc.
2	Covalent bond	Sharing of electrons	CO ₂ , H ₂ O, NH ₃ , etc.
3	Metallic bond	Attraction force between free electron their atomic nuclei	Na, Fe, Mg, etc.

3.2 Ionic Bond

Factors affecting formation of ionic bonding

- i. Ionization energy:- element with low ionization favors the formation of ionic bonding
- ii. Electro affinity:- elements with high electro affinity favors formation of ionic bond
- iii. Lattice energy(U):- is energy released when one mole of ionic solid is formed from its ions
: - it affects strength, melting point and solubility of ionic compounds

Factors affecting lattice energy

- a. Atomic size:- as atomic size decrease lattice energy and melting point increase
- b. Magnitude of the charge:- as magnitude of charge increase lattice energy increase

- U is calculated by using Hess's law through the Born-Haber cycle
- As U increases solubility of ionic compound in water decreases

Ionic compound that violate the Octet rule

No	Ionic compound	example
1	Below Octet	Compounds of Be and Li such as BeF_2 , Li_2O , etc.
2	Above Octet	Cations of transition and post transition metal such as FeCl_3 , AgCl , etc

3.3 Covalent compound that violate the Octet rule

No	Covalent compound	Example	Definition
1	Below Octet	BeCl_2 , AlCl_3 , etc	Covalent compounds have less than 8 electrons around central atom
2	Above Octet	SF_6 , SbF_5 , TeF_6 , AsF_5 , etc	Covalent compounds have greater than 8 electrons, around central atom
3	Odd number of electron	NO_2 , NO , ClO_2 , etc	All covalent compounds with odd number of valence electron

3.3.1 Coordinate covalent bond /dative/donor-acceptor

✓ The shared electrons are obtained from one atom and the other atom has vacant orbital

3.3.2 Polar and non-polar covalent bonds

No	Covalent bonds	Definition
1	Polar	-Formed between atoms with different electronegativities -Are shared unequally -Dipole-moment greater than zero
2	Non polar	-Formed when electrons are shared equally -have no dipole moment

3.3.3 Difference between ionic compound and covalent compound

No	Covalent compound	Ionic compound
1	Have lower boiling point	Have high boiling point
2	Non polar covalent compounds are non-electrolyte	Can conduct electric current
3	Non polar covalent compounds are soluble in non-polar solvent	They are soluble in polar solvent
4	Polar covalent compound are soluble in polar solvent	They are crystalline solid
5	Exist in three states	

3.4 Types of electron pair arrangement

No	No of electron pair	Molecular formula	Molecular geometry	Example
1	2	AX_2	linear	$BeCl_2$, CO_2 , HCN , etc.
2	3	AX_3	Trigonal planar	BF_3 , $HCHO$, NO_3^- , etc.
		AX_2E	V-shaped	SO_2 , O_3 , $PbCl_2$, $SnBr_2$, etc.
3	4	AX_4	tetrahedral	CH_4 , NH_4^+ , SO_4^{2-} , etc
		AX_3E	Trigonal pyramid	NH_3
		AX_2E_2	V-shaped	H_2O
4	5	AX_5	Trigonal bi pyramidal	PCl_5
		AX_4E	See saw	SF_4 , XeO_2F_2 , IF_4^+ , etc
		AX_3E_2	T-shaped	ClF_3 , LCl_3 , BrF_3 , etc
		AX_2E_3	Linear	XeF_2 , I_3^- , IF_2^- , etc
5	6	AX_6	Octahedral	SF_6 , IOF_5 , etc
		AX_5E	Square pyramidal	BrF_5 , TeF_5^- , $XeOF_4$, etc
		AX_4E_2	Square planar	XeF_4 , ICl_4^- , etc

3.5 Intermolecular forces in covalent compound

-It is attractive force between molecules

No	Type of intermolecular force	Definition	example
1	Dipole-dipole force	-Exist between polar molecules -the higher dipole moment the stronger force strength	HCl
2	Hydrogen bonding	-Stronger type of dipole - dipole -formed between that have a hydrogen atom bound to most electronegative	HF , H_2O , CH_3COOH , etc
3	London dispersion force	-exist between non polar molecules -weakest intermolecular force -strength affected by size -as size increases strength of dispersion force increases	$I_2 > Br_2 > Cl_2$ $C_2H_6 > C_3H_8$
4	Ion-dipole force	-occur between ions and polar molecules -responsible for solubility of ionic compounds	$AgCl_2$

3.5.1 Difference between intermolecular force and intramolecular forces

No	Intermolecular forces	Intramolecular forces
1	Affect physical property	Affect chemical property
2	Hold molecules together	Hold atoms together in a molecule
3	Weaker than intermolecular forces	Stronger than intermolecular forces

3.6 Chemical Bonding theories

3.6.1 Valance bond theories (VBT)

✓ A covalent bond is formed when half-filled orbitals overlap

No	Types of covalent bond	σ bond	π bond
1	Single bond	1	0
2	Double bond	1	1
3	Triple bond	1	2

✓ The higher the overlap the stronger is the bond

✓ Sigma bond is stronger than pi bond

Sigma bond	Pi Bond
Overlap end to end, head-to-head, and linear overlap	Overlap side to side or parallel way

3.6.2 Hybridization of orbitals

It is the intermixing of two or more pure atomic orbitals of the same atom

- ✓ Provides useful and convenient method of predicting the shape of molecules
- ✓ does not explain the reason for the shape

Number of electron sets	hybridization
2	Sp
3	sp ²
4	sp ³
5	Sp ³ d
6	Sp ³ d ²

3.6.3 Molecular orbital theory

It is when any number of atomic orbital overlaps; they combine to form unequal number of new orbitals

- ✓ Bonding molecular orbital(σ or π) :
 - molecular orbitals obtained by addition of atomic orbitals
 - Have lower energy than original atomic orbitals
- ✓ Anti-bonding molecular orbitals :
 - molecular orbitals formed by subtraction of atomic orbitals(σ^* or π^*)
 - Have higher energy than original atomic orbitals
- ✓ Molecular orbital theory principle
 - Assign electrons to molecular orbitals ,the Pauli exclusion principle and Hund's rule apply

Rule 1: for simple homonuclear diatomic molecules that have 14 or below electrons.
The order of filling is:

$$(\sigma_{1s})^2 (\sigma_{1s}^*)^2 (\sigma_{2s})^2 (\sigma_{2s}^*)^2 (\pi_{2py})^2 = (\pi_{2pz})^2 (\sigma_{2px})^2 (\pi_{2py}^*)^2 = (\pi_{2pz}^*)^2 (\sigma_{2px}^*)^2$$

Rule 2: for simple homonuclear diatomic molecules that have above 14 electrons .the order of the filling is:

$$(\sigma_{1s})^2 (\sigma_{1s}^*)^2 (\sigma_{2s})^2 (\sigma_{2s}^*)^2 (\sigma_{2px})^2 (\pi_{2py})^2 = (\pi_{2pz})^2 (\pi_{2py}^*)^2 = (\pi_{2pz}^*)^2 (\sigma_{2px}^*)^2$$

- ✓ When two atomic orbitals overlap end to end they form σ molecular orbital

- ✓ When two atomic orbitals overlap in side by side way they form pi molecular orbital
- ✓ Bond order(BO):- $BO = \frac{1}{2}(n_b - n_a)$, where n_b is bonding orbital and n_a is
- ✓ Anti-bonding orbital.
 - As BO increases bond length decreases and bond energy increases
 - It indicates a covalent bond is single, double or triple
 - As the value of bond order increases, the stability of the species increases
 - $B.O = 0$, means the given species does not exist
- ✓ Paramagnetic property: -
 - the species is attracted by external magnetic field
 - The species in which all the electrons are paired
- ✓ Diamagnetic property:-
 - the species is repelled (slightly repelled) by magnetic field
 - The species give unpaired electrons

3.7 Types of solids

No	Types of solids	Properties	example
1	Crystalline solid	-well defined forces -melt at definite temperature	-metals -ionic compounds
2	Amorphous solid	-no well-defined shape -soften shape then melt over range of temperature	-paraffin -asphalt -rubber -window glass

3.7.1 Types of crystals

No	Types of crystalline	Interaction forces	Properties	example
1	Ionic crystal	Ionic bond	-hard and brittle -high melting point -transmit electricity in molten or aqueous state Contain ions	KCl MgO NaBr
2	Covalent network crystal	Covalent bond	-contain atoms -low volatility -poor conductor -hard and brittle	-SiO ₂ -SiC -diamond
3	Molecular crystal	Intermolecular force	-contain molecule -low melting point -high volatility -hard and brittle -poor conductors	-solid H ₂ O -solid CO ₂ -sugar -C ₁₂ H ₂₂ O ₁₁
4	Metallic crystal	Metallic bond	-contain atoms -very soft to hard -malleable and ductile -very low to high melting point	-Cu,Fe,Ag,etc

3.8 Questions

Choose the correct answer from the given alternatives

- Which of the following compounds is unlikely to contain covalent bonds?
A. NaCl B. CO₂ C. NO D. H₂O
- Which of the following is not a form of chemical bonding?
A. Covalent bonding C. metallic bonding
B. ionic bonding D. hydrogen bonding
- Three elements A, B, C and D have atomic number Z-2, Z+2, Z-1 and Z+2 respectively if Z is 10 the bond between which pair of elements will be covalent?
A. A and B B. D and B C. B and C D. A and B
- All of the following there are characteristics of most covalent compounds except?
A. Low melting point C. high electrical conductivity
B. Solubility in inorganic solvent D. low boiling point
- How many ionic bonds result for the formation of magnesium chloride?
A. 1 B. 3 C. 4 D. 2
- Which of the following compounds would be expected to have the highest melting point?
A. MgF₂ B. MgBr₂ C. MgCl₂ D. MgI₂
- Which of the following ionic compounds has the highest lattice energy?
A. CsBr B. LiF C. NaF D. KI
- The lattice energy of cesium chloride, CsCl from the given data below is
Given: $\Delta H_{\text{sub}} = 44\text{kJ/mol}$
 $\Delta H_{\text{Ea}}(\text{Cl}) = -349\text{kJ/mol}$
 $\Delta H_{\text{IE}}(\text{Cs}) = 376\text{kJ/mol}$
 $\Delta H_{\text{diss}}(\text{Cl}_2) = 243\text{kJ/mol}$
 $\Delta H_{\text{f}}\text{CsCl} = -423\text{kJ/mol}$
A. -649.5kJ/mol B. 700.4kJ/mol C. 540.1kJ/mol D. 600kJ/mol
- Which of the following ionic compounds does not obey the Octate rule?
A. -649.5kJ/mol B. 700.4kJ/mol C. 540.1kJ/mol D. 600kJ/mol

10. How many bonding pairs and lone pairs, respectively does on the ion I_3^- have?
- A. 2,8 B. 4,4 C. 3,7 D. 2,9
11. What does the correct Lewis structure for PH_4^+ ion show?
- A. 5 bonds C. no unshared electron
B. 10 shared electrons D. 9 valance electron
12. Which of the following molecules does not have coordinate covalent bonds?
- A. BF_3NH_3 B. POBr_3 C. CO D. NH_3
13. Which of the following compounds have resonance structure?
- A. CH_4 B. SO_3 C. SF_6 D. CCl_4
14. Which molecule has Lewis structure that does not obey the Octate rule?
- A. NO_2 B. NH_4^+ C. SO_2 D. HCl
15. Which bond between the following groups of elements will be the most polar?
- A. H-I B. H-F C. H-Br D. H-Cl
16. Dibromo ethylene $\text{C}_2\text{H}_2\text{Cl}_2$, has the following Lewis structures. Which one has a dipole moment of zero?
- | | | |
|-----------------------|----------------------------|--------------------------------|
| 1,1-dichloro ethane I | Cis-1,2-dichloro ethane II | Trance-1,2-dichloro ethane III |
|-----------------------|----------------------------|--------------------------------|
- A. II B. I and II C. III D. I and III
17. Which of the following molecules is polar?
- A. Cs_2 B. BCl_3 C. CH_2Cl_2 D. PH_3
18. Which one has lower boiling point?
- A. ICl B. I_2 C. HCl D. H_2O
19. All of the following when added to water will produce electric solution?
- A. NaCl B. O_2 C. $\text{Ca}(\text{OH})_2$ D. HCl
20. Which of the following molecule does not have a tetrahedral atom?
- A. CCl_4 B. NH_4^+ C. CH_4 D. PH_4^-
21. Arsenic (As) is a group V element, what will be the molecular geometry and number of lone pair electrons respectively that exist in the ion $(\text{AsCl}_5)^{2-}$?
- A. Square pyramidal,1 B. square planar, 1 C. See Saw, 2 D. linear, 2

22. Which of the following is not probable shape of a molecule?

- A. Tetrahedral, SnCl_4
- B. Trigonal by pyramidal, SbCl_5
- C. bent, CO_2
- D. linear, I_3^-

23. Which of the following explains, why at room temperature, Cl_2 is a gas and Br_2 is liquid?

- A. London dispersion force
- B. Covalent bonding
- C. dipole – dipole force
- D. hydrogen bonding

24. Of the following I_2 , H_2O , Ne and NH_3 , which is likely to have the smallest London dispersion forces?

- A. I_2
- B. Ne
- C. NH_3
- D. H_2O

25. Which of the following molecules dominate in ion-dipole forces?

- A. CO_2
- B. CH_3OH
- C. IF
- D. AgNO_3

26. The correct increasing order of CsF , N_2 , $\text{C}_2\text{H}_5\text{OH}$, CH_3Cl ?

- A. $\text{N}_2 < \text{CH}_3\text{Cl} < \text{C}_2\text{H}_5\text{OH}$
- B. $\text{CsF} < \text{CH}_3\text{Cl} < \text{N}_2 < \text{C}_2\text{H}_5\text{OH}$
- C. $\text{N}_2 < \text{C}_2\text{H}_5\text{OH} < \text{CH}_3\text{Cl} < \text{CsF}$
- D. $\text{CsF} < \text{N}_2 < \text{CH}_3\text{Cl} < \text{C}_2\text{H}_5\text{OH}$

27. Which of the following transmit electricity in aqueous state?

- A. Sodium
- B. Sulphur
- C. potassium chloride
- D. sugar

28. Which one of the following element has greater metallic character?

- A. Na
- B. Cs
- C. Li
- D. K

29. Which molecules listed below three sigma and one pi (π) bonds?

- A. CH_4
- B. N_2H_2
- C. C_3H_6
- D. O_2

30. How many atomic orbitals are required for an SP^2 hybridization?

- A. 3
- B. 4
- C. 6
- D. 8

31. What is the hybridization of the oxygen atom attached to hydrogen in acetic acid (CH_3COOH)?

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

32. Propanal has a molecular formula of $\text{CH}_3\text{CH}_2\text{COH}$. what are the possible hybridizations that exist in the molecule?

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

33. What is the type of orbital overlap for the formation of C_2H_4 ?

- A. P overlap
- B. S overlap
- C. sp^3
- D. Sp^2

34. According to the molecular orbital theory which one of the following is diamagnetic?

- A. C_2 B. C_2^+ C. C_2^- D. O_2

35. What would happen to the N_2 molecule up on ionization to N_2^-

- A. The bond energy will decrease and the bond length will decrease
B. The bond length will increase and the bond energy will increase
C. The bond length will decrease and the bond energy will increase
D. The bond length will increase and the bond energy will decrease

36. Which of the following is the most stable?

- A. O_2^- B. O_2 C. N_2 D. O_2^+

37. Which of the following is a characteristics of bonding molecular orbital?

- A. It is a molecular orbital with a high probability of finding electrons in the region between bonded atoms
B. It is a molecular orbital with a high probability of finding electrons away from the region between bonded atoms
C. It is a molecular orbital with high probability of finding electrons in the region and away from the region between bonded atoms
D. It has no enough electrons

38. Which of the following is covalent crystal?

- A. SiO_2 B. KCl C. sugar D. Au

39. Which of the following is attractive force in the molecular crystal?

- A. Ionic bond B. Metallic bond C. covalent bond D. intermolecular force

40. Which of the following is not crystalline solid?

- A. Mg B. Asphalt C. sugar D. $CaCl_2$

3.9 Answer

1. A	2. D	3. B	4. C	5. D	6. A	7. B	8. A
9. C	10. D	11. C	12. D	13. B	14. A	15. B	16. C
17. D	18. B	19. B	20. D	21. A	22. C	23. A	24. B
25. D	26. A	27. C	28. B	29. B	30. A	31. C	32. D
33. B	34. A	35. D	36. C	37. A	38. C	39. D	40. B

4 UNIT FOUR CHEMICAL KINETICS FOR GRADE 11th

Unit Outcomes

At the end of this unit, you should be able to:

- ✓ explain what is meant by reaction rate and perform activities that determine it;
- ✓ demonstrate an understanding of the dependence of reaction rate on the nature of reactants, the surface area of solid reactants, the concentration of the reactants, temperature of a system and the presence and nature of catalysts;
- ✓ determine reaction rates, using experimental data and calculations;
- ✓ describe how reaction rate theories (collision theory and transition-state theory) explain changes in reaction rates;
- ✓ determine the rate laws and order of reactions from data on initial concentrations and reaction rates;
- ✓ understand that most reactions occur as a series of elementary steps in reaction mechanisms;
- ✓ appreciate the importance of chemical kinetics in industry and in living organisms;

4.1 Chemical Kinetics

Studies the rate at which a chemical process occurs. Besides information about the speed at which reactions occur, kinetics also sheds light on the reaction mechanism (exactly how the reaction occurs).

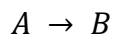
4.2 Reaction Rates

Rates of reactions can be determined by monitoring the change in concentration of either reactants or products as a function of time t .

For the reaction $A \rightarrow B$ there are two ways of measuring rate:

1. The speed at which the reactants disappear
 2. The speed at which the products appear
- ✓ Reversible reactions: as products accumulate, they can begin to turn back into reactants.
 - Early on, the rate will depend on only the amount of reactants present. We want to measure the reactants as soon as they are mixed.
 - A general way of measuring the rate of the reaction is in terms of change in concentration per unit time...

Instantaneous rate = $\Delta[A]/\Delta t$ limits to $d[A]/dt$, Most Common Units... Rate = M/s Where Molarity (M) Reaction rate is the change in the concentration of a reactant or a product with time (M/s).



$$\text{rate} = -\frac{\Delta[A]}{\Delta[t]}; \quad \Delta[A] = \text{change in concentration in A overtime period of } \Delta t$$

$$\text{rate} = \frac{\Delta[B]}{\Delta t}; \quad \Delta[B] = \text{change in concentration in B over time period of } \Delta t$$

Reaction Rates and Stoichiometry

To generalize, for the reaction; $aA + bB \rightarrow cC + dD$

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

4.3 Factors Affecting Reaction Rate Constants

Rate of a chemical reaction is influenced by the following factors (i) Temperature (ii) Concentration of the reactants (iii) Nature of reactants (iv) Catalysts (v) Temperature

1. Temperature:

At higher temperatures, reactant molecules have more kinetic energy, move faster, and collide more often and with greater energy

- Collision Theory: When two chemicals react, their molecules have to collide with each other (in a particular orientation) with sufficient energy for the reaction to take place.
- Kinetic Theory: Increasing temperature means the molecules move faster.

2. Concentrations of reactants

- As the concentration of reactants increases, so does the likelihood that reactant molecules will collide

3. Catalysts

- Speed up reactions by lowering activation energy and by providing the alternative path of the chemical reaction catalyst can be homogeneous and heterogeneous

4. Surface area of a solid reactant

- More area for reactants to be in contact
- Surface area of reactants
- Increased surface area of reactant increases the rate of reaction

4.4 Concentration and Rate

Each reaction has its own equation that expresses its rate as a function of the concentrations of the involved species (e.g., reactants, products, catalysts). This is called its Rate Law

Rate Law

In general, rates of reactions increase as concentrations increase since there are more collisions occurring between reactants.

The overall concentration dependence of reaction rate is given in a rate law or rate expression.

For reactions follow simple rate laws: $r = k [A]^m [B]^n$.

where, [A], [B]: reactant concentrations

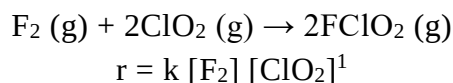
The exponents m and n: reaction order (w.r.t. specific reactant)

The constant k: rate constant

The overall reaction order is the sum of the reaction orders: $m + n$

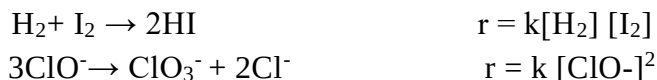
Determination of Rate Laws

- ✓ Rate laws, rate constants, and orders are determined experimentally.
- ✓ The order of a reactant is NOT generally related to its stoichiometric coefficient in a balanced chemical equation



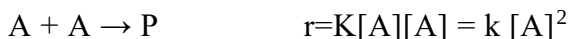
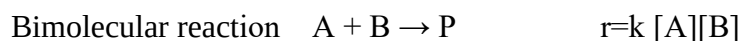
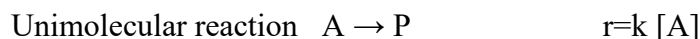
Expression of Rate Laws

Reactions with simple rate laws



Elementary reactions

Always follow simple rate laws



4.5 Order of Reactions

A reaction is zero order in a reactant if the change in concentration of that reactant produces no effect

A reaction is 1st order if doubling the concentration causes the rate to double

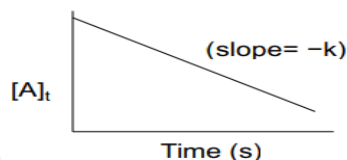
A reaction is 2nd order if doubling the concentration causes a quadruple increase in rate.

Zero-Order reaction

Reaction rate = $-\frac{d[\text{A}]}{dt} = k$ (constant)

$$\text{So... } [\text{A}]_t = -kt + [\text{A}]$$

This equation has the general form for a straight line, $y=mx+b$, so a plot of $[A]_t$ vs. t is a straight line with slope $(-k)$ and intercept $[A]_0$



First-Order reaction

For 1st order reactions: $\ln[A]_t = -kt + \ln[A]_0$

where $[A]_t$ = concentration of $[A]$ after some time, t ;

k = reaction rate constant in units of s^{-1}

t = time in seconds $[A]_0$ = initial concentration of A

This equation has the general form for a straight line, $y=mx+b$, so a plot of $\ln[A]_t$ vs. t is a straight line with slope $(-k)$ and intercept $\ln[A]_0$.



Second-Order reaction

For 2nd order reactions: $1/[A]_t = kt + 1/[A]_0$

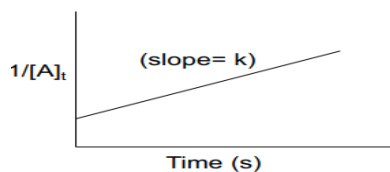
where $[A]_t$ = concentration of $[A]$ after some time, t

k = reaction rate constant in units of $M^{-1}s^{-1}$

t = time in seconds

$[A]_0$ = initial concentration of A

This equation has the general form for a straight line, $y=mx+b$, so a plot of $1/[A]_t$ vs. t is a straight line with slope (k) and intercept of $1/[A]_0$.



Half-Life

Half-life is the time taken for the concentration of a reactant to drop to half its original value. Substitute into integrated rate laws: $t_{1/2}$ is the time t

$$\text{Zero order reaction, } t_{1/2} = \frac{[A]_0}{2K}$$

Half-Life - 1st order

For a first-order process, set $[A]_t = 1/2 [A]_0$ in integrated rate equation

$$\ln[A] = \ln[A]_0 - Kt$$

$$t_{1/2} = \frac{0.693}{K}$$

Example

What is the half-life of N_2O_5 if it decomposes with a rate constant of $5.7 \times 10^{-4} \text{ s}^{-1}$?

$$t_{1/2} = \ln 2 / k = 0.693 / 5.7 \times 10^{-4} \text{ s}^{-1} = 1200 \text{ s} = 20 \text{ minutes}$$

Half-Life - 2nd order

For a second-order process, set $[A]_t = 1/2 [A]_0$ in 2nd order equation

$$\frac{1}{[A]} = \frac{1}{[A]_0} + 2Kt$$

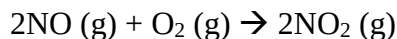
$$t_{1/2} = \frac{1}{K[A]_0}$$

4.6 Reaction Mechanisms

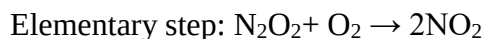
The sequence of events that describes the actual process by which reactants become products is called the reaction mechanism. Reactions may occur all at once or through several discrete steps.

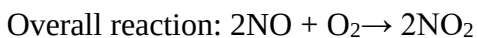
Each of these processes is known as an elementary reaction or elementary process. The overall progress of a chemical reaction can be represented at the molecular level by a series of simple elementary steps or elementary reactions.

The sequence of elementary steps that leads to product formation is the reaction mechanism.



N_2O_2 is detected during the reaction





4.7 Rate Laws and Rate Determining Steps

The sum of the elementary steps must give the overall balanced equation for the reaction. The rate-determining step should predict the same rate law that is determined experimentally.

The rate-determining step is the slowest step in the sequence of steps leading to product formation. Up until now, we have only been concerned with the reactants and products. Now we will examine what path the reactants took in order to become the products.

The reaction mechanism gives the path of the reaction.

Mechanisms provide a very detailed picture of which bonds are broken and formed during the course of a reaction

Elementary Steps & Molecularity

- ✓ Elementary step: any process that occurs in a single step.
- ✓ Molecularity: number of molecules present in an elementary step

Unimolecular: one molecule in the elementary step,

Bimolecular: two molecules in the elementary step,

Termolecular: three molecules in the elementary step.

Rate Laws of Elementary Steps

Since this process occurs in one single step, the stoichiometry can be used to determine the rate law!

Law of Mass Action: The rate of a simple (one-step) reaction is directly proportional to the concentration of the reacting substances.

4.8 Activation Energy

There is a minimum amount of energy required for a reaction: the activation energy, E_a . Just as a ball cannot get over a hill if it does not roll up the hill with enough energy, a reaction cannot occur unless the molecules possess sufficient energy to get over the activation energy barrier.

Molecules must possess a minimum amount of energy to react. Why?

In order to form products, bonds must be broken in the reactants. Bond breakage requires energy. Molecules moving too slowly, with too little kinetic energy, do not react when they collide.

- Activation energy, E_a , is the minimum energy required to initiate a chemical reaction
 - ✓ The change in energy ΔE for the reaction is the difference in energy between CH_3NC (reactant) and CH_3CN (product)

- ✓ The activation energy E_a is the difference in energy between reactants, CH_3NC , and the transition state.
- ✓ The rate depends on E_a . If the “hill” is taller, the reaction rate is slower. If the “hill” is shorter the rate is faster
- ✓ Notice that if a forward reaction is exothermic... ($\text{CH}_3\text{NC} \rightarrow \text{CH}_3\text{CN}$), then the reverse reaction is endothermic... ($\text{CH}_3\text{CN} \rightarrow \text{CH}_3\text{NC}$).
- ✓ The methyl isonitrile molecule needs to gain enough energy to overcome the activation energy barrier.
- ✓ From kinetic molecular theory, we know that as temperature increases, the total kinetic energy increases and the number of molecules with energy greater than E_a increases.
- ✓ So as long as the temperature is high enough, the reaction can make it “over the hill” and proceed.

4.9 EXERCISE

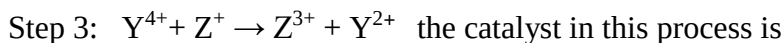
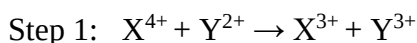
Given data

Trial	[W]	[X]	[Y]	Rate
1	0.001	0.001	0.00.	2.1×10^{-4}
2	0.002	0.001	0.001	2.1×10^{-4}
3	0.001	0.003	0.001	6.3×10^{-4}
4	0.002	0.001	0.0005	5.25×10^{-4}

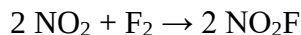
- The initial-rate data in the table above were obtained for the reaction represented below. What is the experimental rate law for the reaction? $\text{A} + \text{B} \rightleftharpoons 2 \text{C}$
 - rate = $k[\text{A}][\text{B}]$
 - rate = $k[\text{A}][\text{B}]^2$
 - rate = $k[\text{A}][\text{B}]^3$
 - rate = $k[\text{A}]^2 [\text{B}]^2$
- Rate = $k[\text{A}]^2$. For the reaction for which rate law is given above, a plot of which of the following is a straight line?
 - $\frac{1}{[\text{A}]}$ versus time
 - $[\text{A}]$ versus 1/ time
 - $[\text{A}]$ versus time
 - $\ln [\text{A}]$ versus time
- $2 \text{X (g)} + \text{Y (g)} \rightleftharpoons 2 \text{Z (g)}$

When the concentration of substance Y in the reaction above is doubled, all other factors being held constant, it is found that the rate of the reaction quadruples. The most probable explanation for this observation is that.

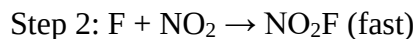
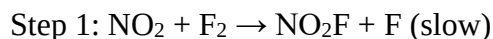
- A. substance X is not involved in any of the steps of the mechanism
 - B. substance Y is involved in the rate-determining step of the mechanism but is not involved in subsequent steps.
 - C. substance Y is probably a catalyst and thus speeds up the rate of the reaction.
 - D. the order of the reaction with respect to Y is 2
4. The proposed steps for a catalyzed reaction between X^{4+} and Z^+ are represented below.



- A. X^{4+}
 - B. X^{3+}
 - C. Y^{4+}
 - D. Y^{2+}
5. The balanced equation for the reaction of nitrogen dioxide and fluorine is



The proposed mechanism is



Which of the following are correct?

- I. The mechanism supports an experimentally determined rate law of $\text{rate} = k[\text{NO}_2]^2[\text{F}_2]$
 - II. F is an intermediate
 - III. The reaction is first order in F_2
- A. I only
 - B. I and II only
 - C. I and III only
 - D. II and III only
6. Each of the following is true about heterogeneous catalyst EXCEPT
- A. Its presence changes the rate of chemical change.
 - B. It does not undergo a permanent change.
 - C. It is in the same phase as the reacting particles.
 - D. Its presence lowers the activation energy of the overall reaction.
 - E. Its presence decreases the potential energy of the activated complex
7. The Reaction below should be used for the next 3 questions



The oxidation of iodide ions by iodate ions in acidic aqueous solution occurs according to the Stoichiometry shown above. The experimentally determined rate law of the reaction is

$$\text{rate} = k[\text{IO}_3^-][\text{I}^-]$$

What is the order with respect to I^-

- A. 1 B. 2 C. 3 D. 5
8. $\text{A(g)} + \text{B(g)} \rightarrow \text{C(g)}$
All of the following apply to the reaction above as it is carried out at constant temperature in a sealed rigid container EXCEPT
- A. The total pressure decreases. C. The entropy of the system decreases
B. The rate of the reaction decreases. D. The number of molecules of C decreases.
9. Consider the hypothetical reaction. $\text{X(g)} + 2\text{Y(g)} \rightarrow \text{XY}_2\text{(g)}$ $\frac{\Delta[\text{Y}]}{t} = -5.0 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$.
What is the rate of formation of $\text{XY}_2\text{(g)}$?
- A. $-5.0 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$ C. $1.0 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$
B. $-2.5 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$ D. $2.5 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$
10. Consider the hypothetical reaction below taking place in a syringe at constant temperature. The reaction is started by placing A(g) and B(g) in the syringe. No chemical equilibrium is established. $\text{A(g)} + 2\text{B(g)} \rightarrow \text{AB}_2\text{(g)}$. Which change causes an immediate increase in the partial pressure B(g) ?
- I. addition of B(g) at constant volume
II. decrease in the volume of the reaction vessel
III. addition of $\text{AB}_2\text{(g)}$ at constant volume
- A. I only B. II only C. I and II only D. II and III only
11. Consider the reaction and its rate law given below for the next three questions
 $2\text{A(g)} + \text{B(g)} \rightarrow \text{C}$ $\text{rate} = k[\text{A}]^2[\text{B}]$
At the beginning of one trial of this reaction $[\text{A}] = 4.0$ and $[\text{B}] = 1.0$. The rate of formation of C was $0.048 \text{ mole L}^{-1} \text{ sec}^{-1}$. The numerical value of k , the rate constant, for this reaction is closest to
- A. 3×10^2 B. 4×100 C. 8×10^{-1} D. 1×10^{-2} E. 3×10^{-3}
12. Which is the label for k , the rate constant?
- A. $\text{mol}^2 \text{ L}^{-2} \text{ sec}^{-1}$ B. $\text{L mol}^{-1} \text{ sec}^{-1}$ C. $\text{L}^2 \text{ mol}^{-2} \text{ sec}^{-1}$ D. $\text{L}^2 \text{ sec mol}^{-1}$

13. When [B] decreases 0.4 M, what will be the value of [A]

- A. 0.8 M B. 1.6 M C. 2.8 M D. 3.4 M

14. Which applies to any reaction mechanism?

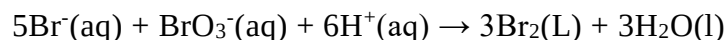
- I. It is a list of steps that produce the overall chemical reaction.
II. It includes only Unimolecular steps.
III. It cannot include reaction intermediates.

- A. I only B. II only C. III only D. I and III only

15. The first order decomposition of some radioactive isotope is 3 days. Approximately what Percentage of the original substance will have decayed after 12 days have passed?

- A. 94% B. 88% C. 50% D. 25% E. 6%

16. Consider the information below for the next questions.



The reaction between bromide ions and bromate ions in acidic water solution occurs according to the equation above. The rate law for this reaction is known to be $\text{rate} = k[\text{Br}^-][\text{BrO}_3^-][\text{H}^+]^2$

The overall order for this reaction is

- A. 2 B. 3 C. 4 D. 6 E. 12

17. Which answer choice correctly matches the integrated rate law shown? $\frac{1}{[\text{A}]t} = kt + \frac{1}{[\text{A}]_0}$ A zero order reaction where a plot of $\frac{1}{[\text{A}]t}$ versus t will be a straight line with slope = k

- A. A first order reaction where a plot of $\frac{1}{[\text{A}]t}$ versus t will be a straight line with slope = k
B. A first order reaction where a plot of $\frac{1}{[\text{A}]t}$ versus t will be a straight line with slope = -k
C. A second order reaction where a plot of $\frac{1}{[\text{A}]t}$ versus t will be a straight line with slope = k
D. A second order reaction where a plot of $\frac{1}{[\text{A}]t}$ versus t will be a straight line with slope = -k

18. Which of the following is true of all catalysts?

- A. They are used up in chemical reactions
B. They are always transition metals
C. They do not take part in chemical reactions.
D. They work by increasing the activation energy of a reaction.
E. They are present at the beginning of a reaction and are unchanged at the end.

19. The reaction of NO_2 with CO is believed to occur according to the following mechanism. What would be the rate expression if the mechanism is correct?



- A. $\text{rate} = k[\text{NO}_2]^2$ B. $\text{rate} = k[\text{NO}_2][\text{CO}]$ C. $\text{rate} = k[\text{NO}_2]^2[\text{CO}]$ D. $\text{rate} = k[\text{NO}_2]^3$
20. If 75% of a sample of pure ^3_1H decays in 24.6 years, what is the half-life of ^3_1H
- A. 24.6 years B. 18.4 years C. 12.3 years D. 6.15 years E. 3.07 years
21. For a first order reaction that has a half-life of 69 s at 80°C , the value of the rate constant, k , is closest to?
- A. 0.01 s^{-1} B. 0.1 s^{-1} C. 1 s^{-1} D. 10 s^{-1} E. 100 s^{-1}
22. If the half-life of a reaction is independent of concentration, what is the order of the reaction?
- A. zero
B. first
C. second
D. half-life is unrelated to the order of the reaction
E. unable to be determined without knowing starting concentrations
23. A particular nuclear decay has a rate constant of 0.0034 min^{-1} . What is the half-life?
- A. 3.3 hours B. 1.6 hours C. 0.0012 min D. 0.0024 min E. 0.0050 min
24. Which of the following does not affect the rate of chemical reaction?
- A. enthalpy of reaction C. Temperature
B. surface area D. concentration of reactant
25. Consider the reaction $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$
For which the rate law is $\text{rate} = k[\text{NO}]^2[\text{O}_2]$, if this reaction takes place in a sealed vessel and the partial pressure of nitric oxide is double. What would this have on the rate of the reaction?
- A. The reaction rate increase by a factor of four
B. The reaction rate increase by a factor of three
C. The reaction rate increase by a factor of eight
D. The reaction rate increase by a factor of two

4.10 ANSWERS

1. B	2. A	3. D	4. D	5. D	6. C	7. A	8. D
9. D	10. C	11. E	12. C	13. C	14. D	15. C	16. C
17. D	18. E	19. A	20. C	21. A	22. B	23. A	24. A
25. A							

5 UNIT FIVE CHEMICAL EQUILIBRIUM AND PHASE EQUILIBRIUM FOR GRADE 11th

Unit Outcomes

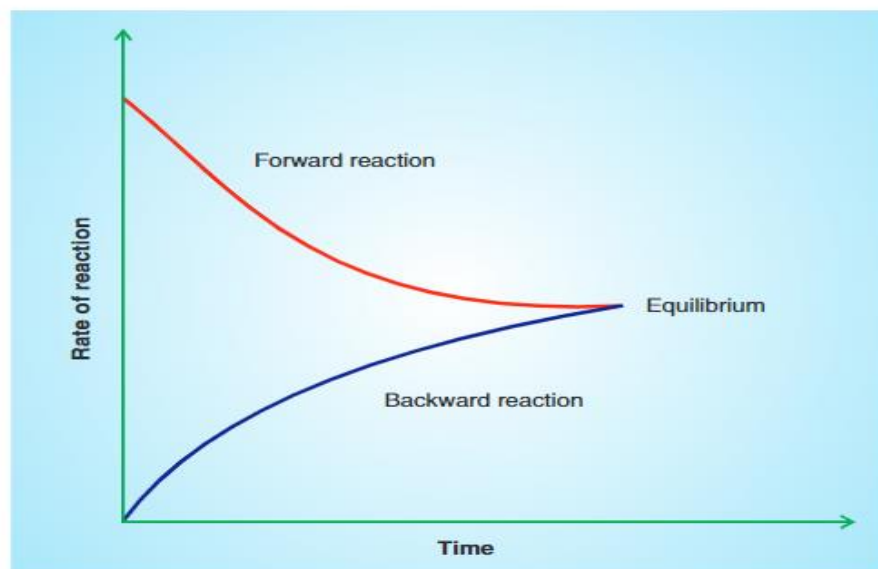
At the end of this unit, you should be able to:

- ✓ understand how equilibrium is established;
- ✓ explain characteristics of dynamic equilibrium;
- ✓ state the law of mass action, and write an expression for equilibrium constants, K_c and K_p , from a given chemical reaction;
- ✓ apply the law of mass action to calculate K_c , and K_p , the concentration and pressure of substances in equilibrium;
- ✓ understand how the reaction quotient is used to indicate the position of equilibrium;
- ✓ state Le Chatelier's principle and use it to predict and explain the effects of changes in temperature, pressure, concentration, and presence of catalyst on a reaction;

- ✓ perform an activity to demonstrate the effects of changes in concentration on the position of equilibrium and to determine K_c and K_p values;

5.1 CHEMICAL EQUILIBRIUM

- ✓ A reversible reaction is a chemical reaction that results in an equilibrium mixture of reactants and products.
- ✓ Chemical reactions which proceed in only one direction are known as irreversible reactions. When the equation for irreversible reaction is written, a single arrow ($\rightarrow$) is used indicating that the reaction can proceed in only one direction. While writing reversible reaction, double arrow ($\rightleftharpoons$) is used indicating that the reaction can go from reactant to product side and also from the product to the reactant side.
- ✓ The equilibrium is dynamic in nature which means that although at the macroscopic level the properties do not change with time, however, at the molecular level the two processes are still taking place.
- ✓ So the equilibrium is attained when the rate of two opposing processes become equal.
- ✓ An equilibrium involving a chemical reaction is known as chemical equilibrium.
- ✓ 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.



Change in the rate of reaction with time for the forward and reverse reactions.

Characteristics of Chemical Equilibrium

- ✓ The reaction is reversible in nature.
- ✓ The reaction takes place in a closed vessel, that is, during the reaction; Reactants and products are neither added nor removed from the reaction vessel.

- ✓ The rate of forward reaction is equal to the rate of reverse reaction.
- ✓ All the reactants and products are present at equilibrium.
- ✓ The macroscopic properties like temperature, pressure, volume and Concentration do not change with time.
- ✓ The equilibrium is dynamic in nature.
- ✓ The state of equilibrium can be obtained from either side.
- ✓ The concentrations of reactants and products are generally not equal.

5.2 Law of Mass Action

The general equation that represents a reversible reaction is: $aA + bB \leftrightarrow mM + nN$

For the forward reaction A and B are reactants and M and N are products

At equilibrium, the rate of forward reaction is equal to the rate of reverse reaction.

Therefore,

$$r_f = r_b$$

or

$$k_f [A]^a [B]^b = k_b [M]^m [N]^n$$

k_f and k_b are constant. Therefore, the ratio of k_f to k_b is also a constant.

$$\frac{k_f}{k_b} = \frac{[M]^m [N]^n}{[A]^a [B]^b} = K_{eq}$$

The ratio of k_f and k_b is represented by K_{eq} . Since this constant represents the reaction at equilibrium it is known as equilibrium constant. $[M]$, $[N]$, $[A]$ and $[B]$ are concentrations at equilibrium.

- ✓ $K_{Cf} = \frac{1}{K_{Cr}}$
- ✓ Homogeneous reaction is a reaction in which all the reactants and the products are present in the same physical state.

When all the reactants and the products are in the gaseous state, their concentrations can be written in terms of partial pressures. In such cases the equilibrium constant is denoted by K_p . The expression for K_p is

$$K_p = \frac{(p_M)^m (p_N)^n}{(p_A)^a (p_B)^b}$$

Units of K_c

The unit of K_c of a reaction depends upon the number of moles of the reactants and products involved in the reaction. Therefore, for a general reaction;



the unit of K_c will be:

$$K_c = \frac{[M]^m [N]^n}{[A]^a [B]^b} = \frac{[\text{mol L}^{-1}]^m [\text{mol L}^{-1}]^n}{[\text{mol L}^{-1}]^a [\text{mol L}^{-1}]^b} = (\text{mol L}^{-1})^{(m+n)-(a+b)}$$

Unit of K_p

The unit of K_p is decided by the unit of pressure. If the partial pressure is expressed in atm, then the unit of K_p , will be:

$$K_p = \frac{(p_M)^m (p_N)^n}{(p_A)^a (p_B)^b} = \frac{(\text{atm})^m (\text{atm})^n}{(\text{atm})^a (\text{atm})^b} = (\text{atm})^{(m+n)-(a+b)}$$

If the partial pressure is measured in kPa, then the unit of K_p will be $(\text{kPa})^{(m+n)-(a+b)}$

- ✓ Relationship b/n K_c and K_p , $K_p = K_c(RT)^{\Delta n}$
- ✓ In heterogeneous reactions, the reactants and products are present in more than one physical state.

Position of equilibrium: K_c can have three types of values.

Case 1: When $K_c > 1$

This indicates that the value of numerator is greater than the denominator. So the formation of products is favoured at equilibrium.

Case 2: When $K_c < 1$

This indicates that the value of numerator is less than the denominator. So the formation of products is not favoured at equilibrium.

Case 3: When $K_c = 1$

This indicates that the rate of forward reaction is equal to the rate of backward reaction.

Reaction quotient is the ratio of concentrations of products to the concentrations of reactants raised to the power of their respective coefficients at any stage after the start of the reaction.

Reaction Quotient in Terms of Molarities

When the concentrations are expressed in molarities, the reaction quotient is Q_c .

$$Q_c = \frac{[M]^m [N]^n}{[A]^a [B]^b}$$

At equilibrium,

$$Q_c = \left\{ \frac{[M]^m [N]^n}{[A]^a [B]^b} \right\}_{eq} = K_c$$

Reaction Quotient in Terms of Partial Pressure

When the concentrations are taken in partial pressures, the reaction quotient is Q_p and is given by the expression

$$Q_p = \frac{(p_M)^m (p_N)^n}{(p_A)^a (p_B)^b}$$

At equilibrium,

$$Q_p = \left\{ \frac{(p_M)^m (p_N)^n}{(p_A)^a (p_B)^b} \right\}_{eq} = K_p$$

- ✓ When Q is less than K , then the reaction will proceed in the forward direction and more products will be formed till the equilibrium is reached.
- ✓ When Q is greater than K , then the reaction will proceed in the reverse direction and more reactants will be formed till the equilibrium is reached.
- ✓ When Q is equal to K , then the reaction has attained equilibrium.
- ✓ The equilibrium constant depends upon the following factors:
 1. Temperature at which the experiment is performed.
 2. The form of equations, which describe the equilibrium.

5.3 Factors affecting chemical equilibrium

Le Chatelier's principle" when a system is a state of equilibrium is disturbed, the equilibrium tends to undergo some change to minimize the effect of disturbance".

1. Temperature

- ✓ Increasing temperature shifts the equilibrium toward the endothermic side.
- ✓ Lowering temperature shifts the equilibrium toward the exothermic side.
- ✓ The forward reaction is endothermic, so the reverse reaction must be exothermic.

2. Pressure

- ✓ Increasing pressure shifts the equilibrium toward the smaller number of moles of gaseous molecules.
- ✓ Decreasing pressure shifts the equilibrium toward the larger number of moles of gaseous molecules.
- ✓ If the number of moles of gaseous molecules of reactants and products are equal the effect of pressure does not change.

3. Concentration

- ✓ If the concentrations of reactant increase the equilibrium will shift to the product side.
- ✓ If the concentrations of product increase the equilibrium will shift to the reactant side.

4. Catalyst

- ✓ Catalysts do affect the equilibrium position, but alter the time at which equilibrium is established.

5. Addition of inert gases on the position of equilibrium:-

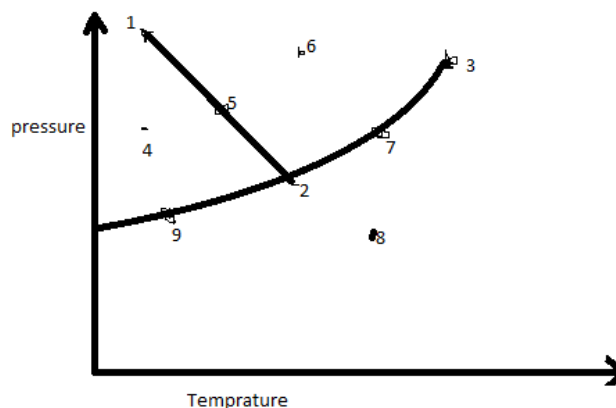
- i. When non-reactive gas is added at constant volume of the system no change on the position of equilibrium.
 - ii. When non-reactive gas is added the volume of the system changes on the position of equilibrium is decide by the stoichiometry of the reaction.
- ✓ $KE \ll \text{attractive force}$, at very low temperature $\Rightarrow$ solid state
 - ✓ $KE \gg \text{attractive force}$, at very high temperature $\Rightarrow$ gaseous state
 - ✓ $KE \approx \text{attractive force}$, at intermediate temperature $\Rightarrow$ liquid state
 - ✓ $F = C - P + 2 \Rightarrow$ Gibbs phase rule: where F is minimum number of variable to describe each component & each phase system, C is chemically distinct quantity present in the system. P is physically and chemically homogeneous but has a discrete boundary & 2 is pure substance.
 - ✓ Triple point is all the three state substances co-exist at equilibrium.
 - ✓ Critical point is the temperature & pressure at which the meniscus b/n the liquid & the vapor disappear.
 - ✓ Critical temperature is the maximum temperature at which the substances co-exist in liquid state.
 - ✓ Critical pressure is the minimum pressure that liquefies the vapors at critical temperature.

5.4 Question

- Which of the following statements describe the conditions for attainment of chemical equilibrium?
 - The rate forward reaction depends up on the concentration of reactants.
 - The concentration of reactants and products does not change.
 - It is also dynamic in nature since the reaction does not stop at equilibrium
 - All are correct.
- What is the unit of K_p expression for the reaction, $N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$
 - atm^2
 - atm^{-2}
 - 1
 - atm
- The conventional equilibrium constant expression (K_c) for the system as described by the equation $NO(g) + \frac{1}{2}Br_2(g) \leftrightarrow NOBr(g)$
 - $[NOBr]/[NO][Br_2]^{1/2}$
 - $[NO][Br_2]^{1/2}$
 - $[NOBr]^2/[NO]^2[Br_2]$
 - $[NOBr]/[NO]^{1/2}[Br_2]$
- For a reaction to shift towards the forward direction, which one of the following condition holds true?
 - $Q_c < K_c$
 - $Q_c = K_c$
 - $Q_c > K_c$
 - $Q_c = K_c = 0$
- If the following reaction is at equilibrium, which of the following changes will shift the equilibrium to the right except
$$2SO_2(g) + O_2(g) \leftrightarrow 2SO_3(g) + 196\text{KJ/mol}$$
 - Increase pressure
 - lowering temperature
 - formation of SO_3
 - increase the volume
- The system $CaO(s) + CO_2(g) \leftrightarrow CaCO_3(s)$ has :-
 - 3 component
 - 1 phase
 - 2 freedom
 - 2 component
- At 400°C , $K_p = 64$ for the reaction $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$ if at equilibrium, the concentration of H_2 & I_2 in a container are 0.20mol^{-1} and 0.50mol^{-1} , respectively, what is the concentration of HI in the mixture?
 - 0.20mol^{-1}
 - 0.50mol^{-1}
 - 2.53mol^{-1}
 - 0.01mol^{-1}
- When the system of the concentration of $A + B \leftrightarrow C + D$ is at equilibrium
 - The sum of the concentration of A & B must be equal.
 - The forward reaction is stopped
 - The reverse reaction is stopped
 - Neither the forward nor the reverse reaction has stopped

9. Consider the following reversible reaction in 3.00 L container $\text{N}_2(\text{g}) + 3\text{H}_2 \leftrightarrow 2\text{NH}_3(\text{g})$, the following amounts are found in equilibrium at 400°C : 0.042mol N_2 , 0.516mol H_2 & 0.0357mol NH_3 ; the value of equilibrium constant is: -
- A. 0.202 B. 1.99 C. 4.94 D. 0.503
10. At 1204°C , K_c for the following reaction is 0.050: $2\text{NO}(\text{g}) \leftrightarrow \text{N}_2(\text{g}) + \text{O}_2(\text{g})$
A mixture of N_2 , O_2 & NO in a vessel at 1204°C has the following concentration $[\text{NO}] = 2.0\text{M}$, $[\text{N}_2] = 0.25\text{M}$ & $[\text{O}_2] = 0.50\text{M}$. Which one of the following statements concerning the reaction quotient, Q_c , is true for the above system?
- A. $Q_c = K_c$; the system is at equilibrium. C. $Q_c > K_c$; more N_2 & O_2 will be produced.
B. $Q_c < K_c$; more N_2 & O_2 will be produced D. $Q_c > K_c$; more NO will be produced.
11. Consider the gas-phase equilibrium system represented by the equation:
 $2\text{H}_2\text{O}(\text{g}) \leftrightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$.
Given that, the forward reaction (the course of "left hand" species to "right hand" species) is endothermic. Which of the following will decrease the equilibrium amount of H_2O ?
- A. Adding a solid phase catalyst C. Increase the temperature at constant pressure
B. Decreasing the volume container D. adding He gas
12. Consider the reversible reaction at equilibrium at 392°C : $2\text{A}(\text{g}) + \text{B}(\text{g}) \leftrightarrow \text{C}(\text{g})$.
The partial pressure are found to be : A: 6.70atm, B: 10.1atm, C: 3.60atm.
The value of K_p for this reaction:-
- A. 7.94×10^{-3} B. 1.46×10^{-3} C. 5.32×10^{-3} D. 5.45×10^{-3}
13. Consider the equilibrium system: $2\text{ICl}(\text{s}) \leftrightarrow \text{I}_2(\text{s}) + \text{Cl}_2(\text{g})$ Which of the following changes will increase the total amount of Cl_2 that can be produced?
- A. Removing some of the $\text{I}_2(\text{s})$ C. Removing the Cl_2 as it is formed
B. Adding more $\text{ICl}(\text{s})$ D. Decreasing the volume of the container
14. $K_c = 0.040$ for the system below at 450°C : $\text{PCl}_5(\text{g}) \leftrightarrow \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$. The K_p value for the reaction at 450°C is:-
- A. 0.40 B. 0.64 C. 2.4 D. 0.052
15. The K_c for the reaction $\text{H}_2\text{S} \leftrightarrow \text{H}_2(\text{g}) + \frac{1}{2}\text{S}_2(\text{g})$ at 1400K is 23.90. The K_c value for the rxn $2\text{H}_2\text{S}(\text{g}) \leftrightarrow 2\text{H}_2(\text{g}) + \text{S}_2(\text{g})$ at 1400K will be?
- A. 571.20 B. 47.80 C. 371.20 D. 59.543

16. The following equilibrium concentrations were observed for the Haber process at 127 °C. Based on the following equation $\frac{1}{2}\text{N}_2(\text{g}) + \frac{3}{2}\text{H}_2(\text{g}) \leftrightarrow \text{NH}_3(\text{g})$. The value of K_c at the same temperature for the reaction is: -
- A. $K_c = K_c$ B. $K_c' = \frac{1}{K_c}$ C. $K_c' = K_c^{1/2}$ D. $K_c' = \sqrt{K_c}$
17. Consider the following hypothetical reaction $\text{A} \leftrightarrow \text{B}$, where $K_c = 5$. Which of the following statement is true?
- A. If 0.1M of A & 0.8M of B are mixed and heated .So $Q > K_c$
 B. If 0.3M of A & 0.2M of B are mixed and heated .So $Q < K_c$
 C. If 0.1M of A & 0.5M of B are mixed and heated .So $Q = K_c$
 D. All are true statement
18. For the reaction $\text{A} \leftrightarrow \text{B}$ & $\text{B} \leftrightarrow \text{C}$ the equilibrium are 800 and 500 respectively. What is the value of the equilibrium constant for the reaction $\text{A} \leftrightarrow \text{C}$?
- A. 4×10^{-5} B. 4×10^4 C. 4×10^5 D. 4×10^{-4}
19. The condition of K_c and K_p are equal. Which of the following is likely to happen when pressure is applied to the equilibrium system?
- A. Equilibrium will shift to the left. C. Little effect on the equilibrium system
 B. No change occurs D. Equilibrium shift to the product side
20. At which point can only the liquid and the gas phases co-exist in the phase diagram of water given below?
- A. 5 B. 7 C. 2 D. 8



5.5 ANSWER

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

6 UNIT SIX HYDRO CARBONS AND DERIVATIVE HYDRO CARBON FOR GRADE 10TH & 11TH

Unit Outcomes

At the end of this unit, you should be able to:

- ✓ write the structural formula and IUPAC names of given alkane, alkene, alkyne, aromatic, aldehyde, ketone, carboxylic acids, and esters;
- ✓ describe some physical and chemical properties of alkane, alkene, alkyne, aromatic, aldehyde, ketone, carboxylic acids, and esters;
- ✓ predict and correctly name the products of organic reactions, including substitution, addition, elimination, esterification, hydrolysis, and oxidation reactions;
- ✓ carry out activities to prepare an alkane, alkene, alkyne, aromatic, aldehyde, ketone, carboxylic acids, and esters;
- ✓ list some important alkane, alkene, alkyne, aromatic, aldehyde, ketone, carboxylic acids, and esters;
- ✓ test for alkane, alkene, alkyne, aromatic, aldehyde, ketone, carboxylic acids and esters; functional groups;
- ✓ understand the structures, properties, uses of alkane, alkene, alkyne, aromatic, aldehyde, ketone, carboxylic acids, and esters;

6.1 HYDRO CARBONS AND DERIVATIVE HYDRO CARBON

- ✓ Friedrich wholer disproved the idea of vital force theory by manufacturing the first synthetic organic compound (urea) in the laboratory.

Organic compounds												
Hydrocarbons			Derivative Hydrocarbons									
Aliphatic		Aromatic	Alcohol	aldehyde		ketone		Carboxylic acid		Ether		Ester
Acyclic	Alicyclic	C_nH_{2n-6} , $n \geq 6$	$C_nH_{2n+2}O$ $n \geq 1$	$C_nH_{2n}O$ $n \geq 1$	$C_nH_{2n}O$ $n \geq 3$		$C_nH_{2n}O_2$ $n \geq 1$		$C_nH_{2n}O$ $n \geq 3$		$C_nH_{2n}O_2$ $n \geq 2$	
Alkane C to C single bond	Cyclo Alkane	Ring form	R-OH	R-(C=O)-H		R-(C=O)- R`		R-(C=O)-OH		R-O-R`		R- (C=O)- OR`
C_nH_{2n+2} $n \geq 1$	C_nH_{2n} $n \geq 3$	Naming suffix→	-ol	-anl		-one		-oic		-ether		-ate
Alkene C to C double bond	Cyclo alkene	Carbon number→	1	2	3	4	5	6	7	8	9	10
C_nH_{2n} $n \geq 2$	C_nH_{2n-2} $n \geq 3$	Prefix →	Meth-	Eth-	Prop-	But-	Pent-	Hex-	Hept-	Oct-	Non-	Dec-
Alkyne Cto C triple bond	Cyclo alkyne	Alkyl group (C_nH_{2n+1}) CH_3 - (methyl), CH_3CH_2 -(ethyl), $CH_3CH_2CH_2$ -(n-propyl), $CH_3(CH-)CH_3$ (isopropyl), $CH_3CH_2CH_2CH_2$ -(butyl), $CH_3(CH-CH_3)CH_3$ (iso- butyl), $CH_3(CH-)CH_2CH_3$ (sec-butyl),(CH_3) ₃ C- (tert-butyl)										
C_nH_{2n-2} $n \geq 2$	C_nH_{2n-4} $n \geq 8$											

6.2 IUPAC name organic compound

- Identify the longest continuous chain.
- Count from the end nearer to the side chain /functional groups.
- If the side chain twice, three times, four times use the prefix di-, tri-, tetra- respectively.
- Number & number separated by comma(,)
- Number & letter separated by hyphens (-)
- Arrange the side chains alphabetically.

- ✓ Bounded of carbon atoms
1C (primary,1°), 2C (secondary,2°), 3C (tertiary,3°) & 4C (quaternary carbonates,4°)
- ✓ The number of isomers of simple alkane use the formula $2^{n-4}+1, n \geq 4$
- ✓ Common isomers of alkene structural ,position & geometrical (cis-trans)
- ✓ Common isomers of alkyne structural & position.

Physical properties of aliphatic hydrocarbons

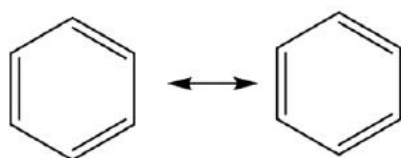
- ✓ Alkane carbon (1-4) gaseous state.
- ✓ Alkene & alkyne (2-4) gaseous state
- ✓ C(5-17) liquid state & C(18) and above are solid
- ✓ Insoluble in water
- ✓ Soluble in non-polar solvent
- ✓ As $c \uparrow$, bpt, mpt & density $\uparrow$ as branched chain $\uparrow$, bpt $\downarrow$

6.3 Chemical properties of aliphatic hydrocarbons

Reaction	Alkane(paraffin's)	Alkene(olefins)	Alkyne
Combination	$C_nH_{2n+2} + \frac{3n+1}{2}O_2 \rightarrow nCO_2 + nH_2O$	$C_nH_{2n+2} + \frac{3n}{2}O_2 \rightarrow nCO_2 + nH_2O$	$C_nH_{2n+2} + 2\frac{3n-1}{2}O_2 \rightarrow nCO_2 + nH_2O$
Elimination	At high temperature to form alkene	-	-
Substitution	Photochemical rxn of bromine & chlorine	-	-
Addition	-	Hydrogenation, halogenation, hydrohalogenation (markovnikov's rule), oxidation by OA of KMnO4 to form dipole	Hydrogenation, halogenation, hydrohalogenation & trimerization at 500-700°C acetylene yields benzene
Preparation	Decarboxylation	Dehydrogenation, dehydrohalogenation alkyl halides with bases	Dehydrohalogenation of vicinal halides with NaOH or KOH
Uses	Fuel, solvent, raw materials plastics, soap, alcohol	Raw material to manufacture polymer	Cutting & welding metals in oxyacetylene torch, vinyl chloride

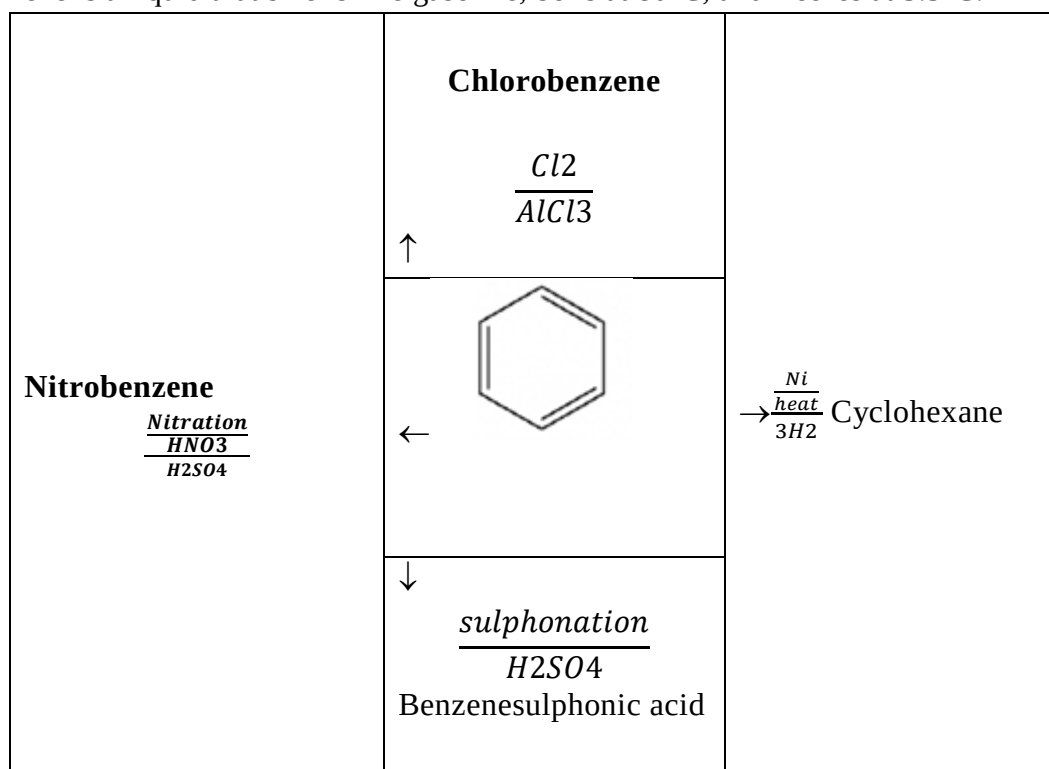
6.4 Aromatic Hydrocarbons

- ✓ Aromatic hydrocarbons are compounds that contain a benzene ring structure.
- ✓ The simplest aromatic compound is benzene (C_6H_6) and it is of great commercial importance, but it also has noteworthy deleterious health effects (see “To Your Health: Benzene and Us”).
- ✓ The formula C_6H_6 seems to indicate that benzene has a high degree of unsaturation. (Hexane, the saturated hydrocarbon with six carbon atoms has the formula C_6H_{14} —eight more hydrogen atoms than benzene.) However, despite the seeming low level of saturation, benzene is rather unreactive. This is due to the resonance structure formed from the alternating double bond structure of the aromatic ring.



The benzene ring structure is stabilized by resonance. One of the few reactions that benzene rings will undergo are substitution reactions.

- ✓ Benzene is a liquid that smells like gasoline, boils at $80^\circ C$, and freezes at $5.5^\circ C$.



6.5 Natural source of Hydrocarbons

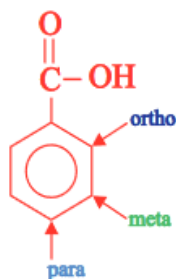
- ✓ Natural gas (methane, ethane, propane & butane)

- ✓ petroleum (alkane, cycloalkane & aromatic hydrocarbons)
- ✓ coal (aromatic hydrocarbons)
- ✓ coke is form of carbons to produce water gas (CO&H₂) and producer gas(CO&N₂)
- ✓ petroleum ether used as solvents/dry cleaning(C5-C7)
- ✓ gasoline(petrol) used as motor fuel(C5-C7)
- ✓ kerosene used as jet engine/house hold fuel(C10-C16)
- ✓ diesel oil used as engine fuel (C15-C16)
- ✓ lubricating oil, grease, petroleum jell, paraffin wax used (>C19)
- ✓ asphalt ,Coke (>C20)

6.6 Alcohol, Aldehyde, Ketone Carboxylic acid & Ester

- ✓ mono hydric alcohol(one-OH),dihydric(2-OH) & trihydric (3-OH)
- ✓ 1° alcohol $\xrightarrow{[O]}$ Aldehyde $\xrightarrow{[O]}$ carboxylic acid
- ✓ 2° alcohol $\xrightarrow{[O]}$ Ketone $\xrightarrow{[O]}$ no rxn
- ✓ 3° alcohol $\xrightarrow{[O]}$ no rxn
- ✓ Carboxylic acid $\xrightarrow{[LiAlH_4]}$ Aldehyde $\xrightarrow{[LiAlH_4]}$ primary alcohol
- ✓ Ketone $\xrightarrow{[LiAlH_4]}$ 2° alcohol
- ✓ Carboxylic acid higher b.pt than alcohol & other hydrocarbons due to the presence of hydrogen bond.
- ✓ Formic acid(ant, Formica), acetic acid(vinegar), butyric acid(butter)& caproic acid(goat)

In common naming system, the name of substituted aromatic acids are written by prefixing the names of substitutes. The position of the substituent is indicated by the prefixes ortho (o-), meta (m-), para (p-) as it is shown in the structure below.



Comparing Physical Properties

Boiling Point:

↑ Carboxylic acid
Alcohols
Aldehydes/Ketones
Ethers
Alkanes

Water Solubility:

↑ Carboxylic acid
Alcohols
Aldehydes/Ketones
Ethers
Alkanes

Name	Molecular weight	Boiling point	Solubility in water
Pentane	72 g/mol	35°C	Insoluble
Diethyl ether	74 g/mol	35°C	Insoluble
Butanal	72 g/mol	76°C	7.1 g / 100 mL H ₂ O
1-Butanol	74 g/mol	118°C	9.1 g / 100 mL H ₂ O
Propanoic acid	74 g/mol	141°C	Infinite

Act
19 Got

6.7 Question

- Which of the following organic molecules has lower boiling point?
A. Alcohols B. Carboxylic acid C. Esters D. Amino acid
- Which alkyl group to give benzoic acid on refluxing with KMnO_4 in alkaline medium?
A. Toluene B. Ethyl benzene C. Iso butyl benzene D. All are correct
- When 'milk' is kept for some time, it becomes sour due to the formation of:-
A. Formic acid B. Acetic acid C. Butyric acid D. Lactic acid
- Which molecular formula indicates 2, 2, 4-trimethyl hexane?
A. C_9H_{20} B. C_9H_{18} C. C_8H_{18} D. C_8H_{16}
- The chlorination of methane to give CCl_4 is an example of: -
A. An electrophilic addition C. A nucleophilic addition
B. A free-radical substitution D. An electrophilic substitution
- The thermal decomposition of alkanes in the absence of air is known as: -
A. Combination B. Oxidation C. Cracking D. Elimination
- Markovnikov's addition of HBr is not applicable to: -
A. Propene B. 1-butene C. 1-pentene D. 2-butene
- Which of the following are correct statements concerning unsaturation hydrocarbons?
A. All unsaturation hydrocarbons are in soluble in water.
B. All unsaturation hydrocarbons are soluble in benzene.
C. All unsaturation hydrocarbons are denser than hydrocarbons.
D. 'A' and 'C' are correct.
- Which of the following is most resist to oxidation?
A. Primary alcohol B. Secondary alcohol C. Dioles D. Tertiary alcohols
- Which of the following products of petroleum are arranged in the correct order of increasing carbon number?

A. Petrol, asphalt, diesel oil, kerosene

B. Asphalt, diesel oil, kerosene, Petrol

C. Petrol, kerosene, diesel oil, asphalt

D. Diesel oil, Petrol, asphalt, kerosene

11. The compound is $C_{37}H_{75}SO_2ONa^+$ is: -

A. detergent

B. fiber

C. Soap

D. Plastic

12. Benzene is essential resonance hybrid of two Kekulé structures. The reason for this is that

A. Individual benzene molecules change back and forth between the two structures.

B. Both structures have equal contribution to the hybrid structure.

C. Half the molecules correspond to the first one structure similar to the second structure

D. At low temperature, benzene can separate into the two structures.

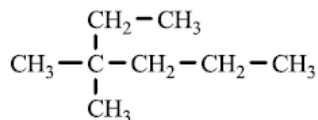
13. Which of the following is mismatched about pharmaceuticals?

A. Analgesics → Paracetamol

C. Disinfectant → Dettol

B. Antiseptic → Aspirin

D. Sulphonamide → Penicillin



14. The IUPAC name of

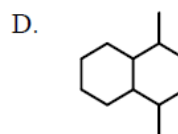
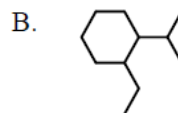
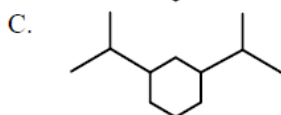
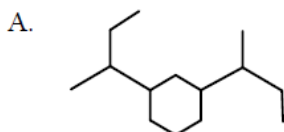
A. 2-ethyl-2-methylpentane

C. 2-ethyl-3-methylhexane

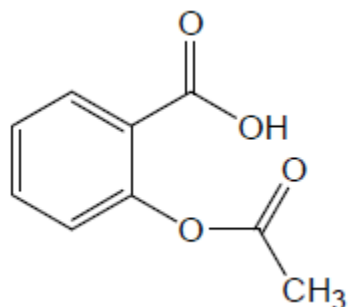
B. 2-methyl-3-ethylpentane

D. 3,3-dimethylhexane

15. The formula of 1,2-diisopropylcyclohexane is: -



16. Identify the functional group present in: -



- A. Alcohol, aromatic, ester
B. Carboxylic acid, aromatic, ester
C. carboxylic acid, aromatic, ether
D. Alcohol, aromatic, ether
17. The common name of the above Q16 is:-
A. Dettol
B. Penicillin
C. Aspirin
D. Amoxicillin
18. Esterification is the reaction of a(an)_____with alcohol.
A. Amide
B. Carboxylic acid
C. Ester
D. Amine
19. The base hydrolysis of an ester is called_____.
A. Esterification
B. Saponification
C. Detergent
D. Fertilizers
20. How many Isomers are there with carbon number 8?
A. 17
B. 15
C. 9
D. 10
21. The special property of carbon is formation of long chain, short chain or ring is known as
A. Isomers
B. Functional group
C. Homologues series
D. Catenation
22. Which of the following does not dissolve Pentane?
A. Benzene
B. Toluene
C. Water
D. Carbon tetrachloride.
23. Which of the following Hydrocarbon exhibit geometric Isomerism?
A. 1-butene
B. 2-butyne
C. 2-butene
D. 1-Pentene
24. For the reaction $\text{CH}_3\text{CH}=\text{CH}_2 + \text{H}_2\text{O} \rightarrow$ _____. The product is;
A. $\text{CH}_3\text{CHOHCH}_3$
B. $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$
C. $\text{CH}_3\text{CH}_2\text{CH}_3$
D. $\text{CH}_2\text{OHCH}_2\text{CH}_3$
25. For Combustion reaction $\text{C}_4\text{H}_8 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$ when it is balanced is:

- A. 13 B. 19 C. 6 D. 11

26. Which is the missing term “X” in the reaction for $\text{CH}_3\text{CH}=\text{CH}_2 + \text{HCl} \rightarrow \text{X}$?

- A. $\text{CH}_3\text{CH}=\text{CHCl}$ B. $\text{CH}_3\text{CH}_2\text{CH}_2\text{Cl}$ C. $\text{CH}_3\text{CH}_2\text{ClCH}_2$ D. $\text{CH}_3\text{CHClCH}_3$

27. What is the process of converting higher Alkanes to smaller?

- A. Polymerization B. Hydrogenation C. Halogenations D. Elimination reaction

28. The name of product for Oxidation of butene in presence of Alkaline/ KMnO_4 is

- A. 1,2-butanediol B. Butane C. Benzene D. Butanol

6.8 ANSWER

1	C	2	B	3	C	4	A	5	B	6	C	7	D
8	D	9	D	10	C	11	A	12	B	13	B	14	D
15	D	16	B	17	C	18	B	19	B	20	A	21	D
22	C	23	C	24	A	25	D	26	D	27	D	28	A

7 UNIT ONE SOLUTIONS FOR GRADE 12th

Unit Outcomes

At the end of this unit, you should be able to:

- ✓ know the types of solutions;
- ✓ understand the process of solution formation, the rate of solution, the heat of solution and solubility;
- ✓ describe the dependence of solubility on the temperature and pressure of the solution;
- ✓ know how to solve problems involving the concentration of solutions and to express the result in various units;
- ✓ describe, the behavior of ionic solute in unsaturated, saturated, and supersaturated solutions using the concept of equilibrium;
- ✓ prepare solutions of a required concentration by dissolving a solute or diluting a concentrated solution;
- ✓ understand the relationship between the colligative properties of solutions and know how to solve problems involving colligative properties of solutions

7.1 INTRODUCTION

A homogeneous mixture is a mixture in which the composition of the mixture is the same. Throughout. i.e., it has no visible boundaries because the components are mixed as individual atoms, ions and molecules. Solutions are homogeneous mixtures of two or more substances. A typical solution consists of at least two substances; a substance in smaller amount called the solute, dissolved in a larger amount of another substance, known as the solvent.

7.2 Suspension and Colloids

Suspensions are heterogeneous mixtures composed of particles that are visible to the naked eye or that can be seen with a magnifying glass. Example, a mixture of mud and water or flour and water.

Colloidal particles are small enough and do not settle down. The particles in a colloid are larger than most simple molecules. Example fog, milk, mud in air.

7.3 Types of Solutions

Type of solution	State of solute	State of solvent	State of solution	Examples
gas-gas solution	Gas	Gas	Gas	Air
gas-liquid solution	Gas	Liquid	Liquid	soda water
solid-liquid solution	Solid	Liquid	Liquid	sugar in water
liquid-liquid solution	Liquid	Liquid	Liquid	ethanol in water
solid-solid solution	Solid	Solid	Solid	Alloy

7.4 The Solution Process

7.4.1 Liquid Solutions and Inter-particle Forces of Attractions

The tendency to mix solute - solvent is affected by the relative strengths of three types of interaction. These are

- ✓ Solvent-solvent interaction
- ✓ Solute-solute interaction
- ✓ Solvent-solute interaction

Strong attraction forces between solute and solvent particles (*solute-solvent interactions*) makes dissolving easier and helps to keep particles in solution. For simplicity.

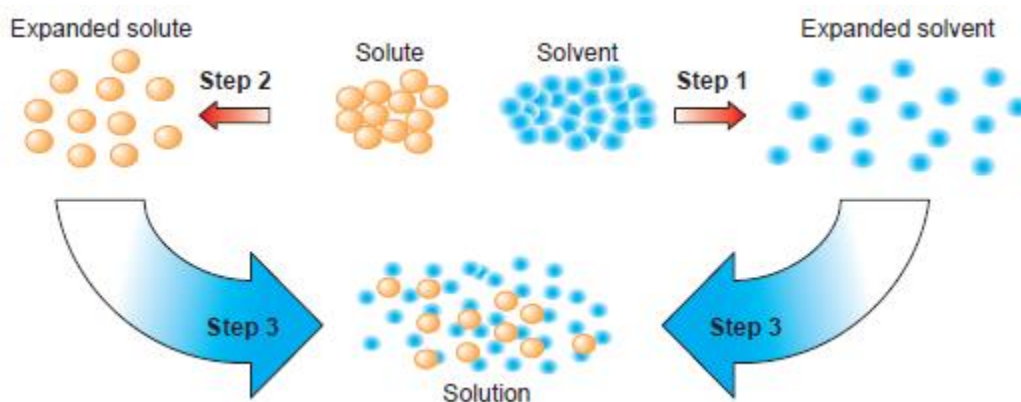


Figure A molecular view of the solution process, portrayed as taking place in three steps.

Step 1 – is the separation of solvent molecules from each other to “make room” for the Incoming solute particles and

Step 2 – is the separation of solute molecules. These two steps require energy input to break attractive intermolecular forces. Therefore, they are endothermic.

Step 3 – the solvent and solute molecules mix. This process can be exothermic or endothermic.

7.4.1.1 Solution of liquids in liquids

The dissolution of solute in a solvent follows the “like dissolves like” principle. “Like dissolves like” means polar dissolves polar and nonpolar dissolves nonpolar.

7.4.1.2 Solutions of Solids in Liquids

Solids composed of polar molecules or ions, are insoluble in non-polar solvents. The weak solute-solvent interaction, compared with the strong attractions within the crystal, is not sufficient to tear apart the lattice. Solids whose crystals are held together by London forces are able to dissolve to appreciable extents in non-polar solvents.

The electrostatic attraction between an ion and a molecule with a dipole is called an ion-dipole attraction. These attractions play an important role in the dissolution of ionic compounds in water.

7.4.2 The Rate of Dissolution

The rate of dissolution is the speed with which a solute goes into solution. This largely depends upon the inter-particle forces and, to a lesser extent, on conditions such as the surface area of the solid solute, and the temperature and the pressure of the system.

Inter-particle force is the extent to which solvent molecules interact with particles of the solute. When the solvent-solute interactions are stronger than those between solute-solute and solvent-solvent particles, the dissolution process becomes easier. Increasing surface area will increase the rate of dissolution because it increases the number of solute particles in contact with the solvent.

7.4.3 Energy Changes in the Solution Process

The amount of energy that is absorbed or released when one mole of a substance enters solution is called *the heat of solution* and is given the symbol $\Delta H^\circ_{\text{Sol}}$, which is the sum of the enthalpies associated with the three steps in the solution process.

- ✓ The separation of solute particles is endothermic.
- ✓ The separation of solvent particles is also endothermic.
- ✓ Each solute particle is surrounded by a layer or “cage” of solvent molecules that is attracted to it by intermolecular forces. This step can be endothermic or exothermic, and is referred to as **solvation**.

The heat of solution $\Delta H^\circ_{\text{Sol}}$

$$\Delta H^\circ_{\text{Sol}} = \Delta H^\circ_1 + \Delta H^\circ_2 + \Delta H^\circ_3 = \Delta H^\circ_{\text{Hydration}} + \Delta H^\circ_{\text{Lattice energy}}$$

Where ΔH°_1 and ΔH°_2 are enthalpies required to separate solute particles and solvent molecules, respectively, and ΔH°_3 is the enthalpy of solvation. If the solute-solvent attraction is stronger than both the solvent-solvent attraction and the solute-solute attraction, the solution process is favorable, or exothermic ($\Delta H^\circ_{\text{Sol}} < 0$). If the solute-solvent interaction is weaker than the solvent-solvent and solute-solute interactions, then the solution process is endothermic ($\Delta H^\circ_{\text{Sol}} > 0$).

7.4.4 Hydration of Ionic Solids in Water

The solubility of ionic compounds in water depends on

- 1) Hydration energy: is the energy released when water molecules surround the ions of the salt. The greater the hydration energy the more likely that the ionic solid will dissolve in water.
- 2) Lattice energy: is the energy required to break one mol e of an ionic solid in to isolated gaseous ions. The lower the lattice energy the more soluble is an ionic solid in water.

7.4.5 Solubility as an Equilibrium Process

A saturated solution is a solution in which the dissolved and undissolved solutes are in dynamic equilibrium.

An unsaturated solution is a solution that can dissolve more solute at a given temperature.

A supersaturated solution, which actually contain larger amounts of solute than in saturated solutions. These solutions can sometimes be prepared by saturating a solution at a high temperature. Supersaturated solutions are unstable.

7.4.6 Ways of Expressing Concentrations of Solution

The concentration of a solution is the amount of solute present in a given quantity of solvent or solution. The concentration of a solution can be expressed either qualitatively or quantitatively. The terms dilute and concentrated are used to describe a solution qualitatively. A solution with a relatively small amount of solute is said to be dilute. One with a large amount of the solutes is said to be concentrated.

Several quantitative expressions of concentration are used in chemistry. Here you are going to learn mass percentage, ppm and ppb, mole fraction, molarity, normality, and molality of a solution.

7.4.7 Mass Percentage, ppm, and ppb of Solute

Percent by mass: The percent by mass, also called percent by weight or weight percent, is the ratio of the mass of a solute to the mass of the solution, multiplied by 100.

$$\text{Percent by mass of solute} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

Parts per million (ppm): When the mass of solute in the solution is very small, a widely used expression is parts per million (ppm).

$$\text{Parts per million (ppm)} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6$$

Or

$$\text{ppm} = \text{Mass percentage of solute} \times 10^4$$

Parts per billion (ppb): A concentration of 1 ppb represents 1 g of solute per billion (10^9) grams of solution.

$$\text{Parts per billion (ppb)} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^9$$

7.4.8 Mole Fraction (X)

The mole fraction (X) of a solute is the ratio of the number of moles of solute divided by the total number of moles of a solution (*moles of solute + moles of solvent*).

$$X_{\text{Solute}} = \frac{n_{\text{solute}}}{n_{\text{solute}} + n_{\text{solvent}}} =$$

7.4.9 Molarity (M)

Molarity or molar concentration is the number of moles of solute in 1 L of solution. Molarity is determined by the equation.

$$\text{Molarity} = \frac{\text{Mass of solute}}{\text{Molar mass of solute} \times \text{Volume (L) of solution}}$$

7.4.10 Normality (N)

The normality of a solution (N) is the number of equivalents of the solute contained in one litre of solution.

$$\text{Normality (N)} = \frac{\text{Number of equivalents of solute}}{\text{Volume (L) of solution}}$$

or

$$\text{Normality (N)} = \frac{\text{Mass of solute}}{\text{Equivalent mass of solute} \times \text{Volume (L) of solution}}$$

Notice that normality is molarity times n . It is always equal to or greater than molarity.

$$N = M \times n$$

Where n is

- Number of ionizable hydrogen ions and hydroxide ions per molecule for acids and bases, respectively.
- Total number of positive charges for salts.
- Total number of electrons lost or gained in redox reactions.

7.4.11 Molality (m)

The molality of a solution is the number of moles of solute per kilogram of solvent contained in a solution.

Note that molal solutions are prepared by measuring masses of solute and solvent, not volumes of solvent or solution.

$$\text{Molality (m)} = \frac{\text{Number of moles of solute}}{\text{Mass (kg) of solvent}}$$

7.4.12 Diluting Solutions

Dilution is the procedure for preparing a less-concentrated solution from a more concentrated One or stock solution. Dilution can be expressed as $M_i V_i = M_f V_f$ or, more generally, $C_i V_i = C_f V_f$

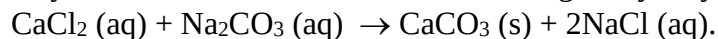
Where C_i is initial concentration, C_f is final concentration, v_i is initial volume and V_f is final volume.

7.5 DESCRIBING REACTIONS IN SOLUTION

Reaction in the solution can be described by using

7.5.1 Molecular Equation

A molecular equation is a chemical equation in which the reactants and the products are written as if they were molecular substances, even though they may actually exist in solution as ions. Example:



The molecular equation closely describes what is actually done in the laboratory or in an industrial process.

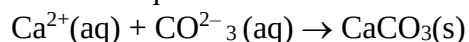
7.5.2 Ionic Equation

Ionic Equation is a chemical equation that contains ions.

Example: $\text{Ca}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq}) + 2\text{Na}^{+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + 2\text{Na}^{+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq})$

The ions appearing on both sides of the equation (Na^{+} and Cl^{-}) are called **spectator ions**, as they do not take part in the reaction and they can be cancelled on both sides to express the essential reaction that occurs.

An ionic equation obtained after removing the spectator ions is **net ionic equation**.



7.6 COLLIGATIVE PROPERTIES OF SOLUTION

Properties that depend on the number of particles or concentration of the solute in the given amount of solvent and not on their nature are called colligative properties. (*Colligative* is a Latin word that means “depending on the collection.”)

There are four important properties of a solution that are affected by the addition of a nonvolatile solute to a solvent. These are:

1. Vapor pressure lowering
2. Boiling point elevation
3. Freezing point depression
4. Osmotic pressure

7.6.1 Vapor Pressure Lowering

If a solute is non-volatile i.e., it does not have a measurable Vapor pressure, the Vapor pressure of its solution is always less than that of the pure solvent. Vapor pressure lowering is the lowering of the Vapor pressure of the solvents that results when a non-volatile solute is added to form a solution. The relationship between solution Vapor pressure and solvent Vapor pressure is expressed by Raoult's law. This law states that “the partial pressure of a solvent over a solution, P_1 , is given by the Vapor pressure of the pure solvent, times the mole fraction of the solvent in the solution, x_1 , at constant temperature”.

$$P_1 = x_1 P_1^{\circ}$$

$$\Delta P = P_1^{\circ} - P_1$$

Where ΔP = Change in Vapor pressure (Vapor pressure lowering)

P_1° = Vapor pressure of pure solvent

P_1 = Vapor pressure of solvent in solution

In a solution containing only one solute, $x_1 = 1 - x_2$, where x_2 is the mole fraction of the solute.

Inserting this into the Raoult's-law equation gives:

$$P_1 = (1 - x_2) P_1^{\circ}$$

$$P_1 = P_1^{\circ} - P_1^{\circ} x_2$$

$$P_1^{\circ} - P_1 = P_1^{\circ} x_2$$

$$\Delta P = P_1^{\circ} x_2$$

7.6.2 Boiling Point Elevation

The boiling point of a liquid is the temperature at which its Vapor pressure equals the atmospheric pressure. If a non-volatile solute is added to a solvent, the solution's boiling point will be increased.

The boiling-point elevation (ΔT_b) is defined as the boiling point of the solution (T_b) minus the boiling point of the pure solvent (T_b°).

$$\Delta T_b = T_b - T_b^\circ$$

The elevation in boiling point (ΔT_b) is directly proportional to the molality of the solution.

That is:

$$\Delta T_b = K_b m$$

Where m is the molality of the solution, and K_b is the molal boiling point elevation constant for the solvent.

7.6.3 Freezing Point Depression

The greater the concentration of the solution, the greater the interference and therefore, the lower the freezing point of the solution.

The freezing point depression (ΔT_f) is defined as the freezing point of the pure solvent (T) minus the freezing point of the solution (T_f).

$$\Delta T_f = T - T_f$$

The freezing point depression (ΔT_f) of solutions of nonelectrolytes has been found to be equal to the molality of the solution times a proportionality constant. This constant is the molal freezing point depression constant, K_f , of the solvent.

$$\Delta T_f = K_f m$$

7.6.4 Osmosis and Osmotic Pressure

Osmosis is the selective passage of solvent molecules through a porous (*semi-permeable*) membrane from a dilute to a more concentrated solution or from the solvent side into the solution side.

The pressure required to prevent osmosis is known as the osmotic pressure, p , of the solution.

Jacobus Henricus Van't Hoff found that the osmotic pressure obeys a law similar to the ideal gas law, $PV = nRT$, where V is the volume of the solution, n is the number of moles of solute, R is the ideal gas constant, and T is the temperature on the Kelvin scale. From the equation, we can write:

$$\pi = p = \left(\frac{n}{V}\right)RT = MRT \text{ where } M \text{ is the molarity of the solution.}$$

Like other colligative properties, osmotic pressure can be used to determine the molar mass of solute. Especially osmotic pressure is used in biological laboratories to determine molecular masses of huge molecules like proteins and nucleic acids.

7.6.5 Colligative Properties of Electrolytic Solutions

The effective number of moles of ions produced by one mole of a solute is expressed in terms of the Van't Hoff factor (i).

$$i = \frac{\text{Measured value for electrolyte solution}}{\text{Expected value for nonelectrolyte solution}}$$

Thus, i is 1 for all solutions of nonelectrolytes, but is greater than 1 for solutions of electrolytes. Therefore, the colligative-property relationship should be modified to account for dissociation of electrolyte solutes. Consequently,

$$\Delta P = i \times 2P^\circ_1$$

$$\Delta T_b = i K_b m$$

$$\Delta T_f = i K_f m$$

$$\pi = i M R T$$

7.7 Questions

- Which of the following is classifying as suspension?
A. Smoke B. cell fluid C. milk of magnesia D. Cloud
- Cans of carbonated soft drinks contains CO_2 gas, under what condition is the solubility of CO_2 gas the greatest
A. low temperature, high pressure C. low temperature, low pressure
B. high temperature, high pressure D. high temperature, low pressure
- A solution in which the dissolved and un dissolved solutes are in dynamic equilibrium is known as: -
A. Unsaturated B. Super saturated C. Saturated D. Diluted
- The rate of dissolution largely depends on: -
A. Temperature of the system C. surface areas of solid solute
B. Inter particle force D. pressure of the system
- A 1000g of sample of hard water was found to 0.01g of CaCl_2 , the concentration of Ca^{2+} in
A. 10 ppm B. 100ppm C. 1000ppm D. 0.01ppm
- How do you prepare 250ml of 0.4N barium hydroxide solution? (At. Mass B= 137, O=16, H=1)
A. By dissolving 8.55g in 500ml C. By dissolving 8.55 in 1000ml
B. By dissolving 16.5 in 1000ml D. By dissolving 4.28 in 250ml
- What is the molarity of a solution of HNO_3 has a density of 1.4g/ml and contain 63%(m/m) HNO_3 (At.m of H=1, N=14, O=16)
A. 1.4M B. 14 M C. 8.82M D. 28M
- The addition of non-volatile solute to the solvent resulting in the lowering of
A. Vapor pressure as well as boiling point
B. Vapor pressure as well as density
C. Boiling point as well as freezing point
D. Vapor pressure as well as freezing point
- Which aqueous solution has the lowest freezing point? (Assume complete dissociation)
A. 0.1m NH_4NO_3 B. 0.1m CaCl_2 C. 0.1m NaCl D. 0.1m $\text{C}_2\text{H}_5\text{OH}$
- A solution of 7.6g of protein in 100cm³ at 300k has 380mm Hg osmotic pressure calculate the molecular mass of the protein($R = 0.082\text{L atm mol}^{-1}$)

- A. 568gmol^{-1} B. 760gmol^{-1} C. 380gmol^{-1} D. 380gmol^{-1}
11. How many ml of water is required to dilute 100ml of 0.5 M HCl to 0.1M HCl?
- A. 500ml B. 400ml C. 100ml D. 300ml
12. A solution is prepared by mixing 1mole of ethanol and 2moles of water at 25°C , what is a total Vapor pressure above the solution? ($p^{\circ}_{\text{wat}} = 44 \text{ torr}$, $p^{\circ}_{\text{eth}} = 23.8 \text{ torr}$)
- A. 33.9torr B. 45torr C. 30.5torr D. 67.8torr
13. The process by which dissolved solute particles becomes surrounded by solvent particle is
- A. Salivation B. Crystallization C. Sublimation D. Dilution
14. The normality of 0.1m solution of $\text{Mg}_3(\text{PO}_4)_2$ is
- A. 6N B. 2N C. 0.3N D. 0.6N
15. Which of the following is an example of liquid – solid solution?
- A. Brass (Cu and Zn) C. Bronze (Cu and Sn)
B. Amalgam (Hg and Ag) D. Steel (Fe and C)
16. What is exothermic step in a solution of salt in water?
- A. Separation of the salt in to ions in gaseous state.
B. Separation of water molecules.
C. Hydration of the ions from the salt.
D. None of the above
17. A 400g of commercial bleaching solution (Berekina) has 20g of NaOCl concentration, what is the concentration of NaOCl in terms of percent?
- A. 20% B. 5% C. 95% D. 2%
18. What is the molality of a solution contain 47.5g of MgCl_2 in 200ml of water?
- A. 2.5m B. 5m C. 0.5m D. 2m
19. Which of the following is a solution?
- A. Milk B. Oil C. Air D. Blood
20. Which of the following organic compound is not readily miscible in carbon tetra chloride, CCl_4 ?
- A. Benzene B. Phenol C. Toluene D. Acetylene

7.8 ANSWERS

1	C	2	A	3	C	4	B	5	A	6	C	7	B
---	---	---	---	---	---	---	---	---	---	---	---	---	---

8	D	9	B	10	D	11	B	12	C	13	A	14	D
15	B	16	C	17	B	18	A	19	C	20	C		

8 UNIT TWO ACID – BASE EQUILIBIRA FOR GRADE 12th

Unit Outcomes

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

- ✓ understand the Arrhenius, Brønsted-Lowry and Lewis concepts of acids and bases;
- ✓ understand the dissociation of water, weak monoprotic and polyprotic acids, and bases;
- ✓ know how to solve equilibrium problems involving concentration of reactants and products, K_a , K_b , pH, and pOH;
- ✓ understand the common-ion effect, buffer solutions, hydrolysis of salts, acid-base indicators, and acid-base titrations;
- ✓ explain how buffering actions affect our daily lives, using some examples;
- ✓ determine the equivalents of acids and base, respectively, that are required to neutralize specific amounts of base and acid;
- ✓ predict, in qualitative terms, whether a solution of a specific salt will be acidic, basic, or neutral;
- ✓ know how to solve problems involving concentration and pH of acid-base titration

8.1 INTRODUCTION

8.2 Acid - Base Equilibrium

- ✓ Arrhenius acid are substances that release H^+ ions and Arrhenius bases release OH^- ion in water solution
- ✓ Brønsted – Lowry acids donate proton and bases accept proton.
- ✓ Lewis acids are electron pair acceptors and Lewis bases are electron pair donors.
- ✓ The Lewis definition is most general and Arrhenius definition is most general and Arrhenius definition has limitations.
- ✓ Every acid has a conjugate base and every base a conjugate acid
- ✓ **Auto / self-ionization** is the interaction of two molecules of the same substances to form positive and negative ions by proton transfer
- ✓ **Amphiprotic species** is a substance that can donate or accept a proton
- ✓ The strength of acid and bases in a aqueous solution can be measure by concentration of H^+ and OH^- , P^H and P^{OH} , per cent ionization and dissociation constant
- ✓ The extent to which a weak acid and a weak base ionizes can be expressed by using acid dissociation constant (K_a) and base dissociation constant (K_b) respectively

8.2.1 Brønsted-Lowry Concept of Acids and Bases

Once an acid has given up a proton, the remaining part can be a proton acceptor and is, now called a conjugate base. On the other hand, when a base accepts a proton, the species formed is called conjugate acid.

Every acid has a conjugate base, and every base has a conjugate acid. Thus, for any conjugate acid-base pair:

- ✓ The conjugate base has one less H and one more minus charge than the acid.
- ✓ The conjugate acid has one more H and one less minus charge than the base.

Table: The conjugate pairs in some Acid-Base Reactions.

Acid	+	Base	$\rightleftharpoons$	Conjugate base	+	Conjugate acid
HF	+	H ₂ O	$\rightleftharpoons$	F ⁻	+	H ₃ O ⁺
HCOOH	+	CN ⁻	$\rightleftharpoons$	HCOO ⁻	+	HCN
NH ₄ ⁺	+	CO ₃ ²⁻	$\rightleftharpoons$	NH ₃	+	HCO ₃ ⁻
H ₂ PO ₄ ⁻	+	OH ⁻	$\rightleftharpoons$	HPO ₄ ²⁻	+	H ₂ O
H ₂ SO ₄	+	N ₂ H ₅ ⁺	$\rightleftharpoons$	HSO ₄ ⁻	+	N ₂ H ₆ ²⁺
HPO ₄ ²⁻	+	SO ₃ ²⁻	$\rightleftharpoons$	PO ₄ ³⁻	+	HSO ₃ ⁻

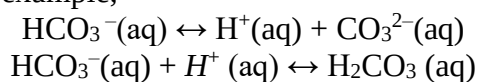
The stronger the acid, the weaker is its conjugate base. Similarly, the stronger the base, the weaker is its conjugate acid. For example, HCl is a strong acid, and its conjugate base Cl⁻, is a weak base. Acetic acid, CH₃COOH, is a weak acid, and its conjugate base, CH₃COO⁻, is a strong base.

8.2.2 Auto-ionization of Substances

The auto-ionization (self-ionization) of a substance involves the transfer of a proton from one molecule of the substance to another molecule of the same substance. For example, H₂O, NH₃.

Amphiprotic Species

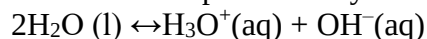
Many molecules and ions gain or lose a proton, under appropriate conditions. Such species are said to be amphiprotic. For example,



In other words, amphiprotic species are species that can act as both an acid and a base.

8.3 Ionization of Water

The self-ionization of water can be represented by the following equation:



The value of K_c , in this case K_w , changes as temperature changes. In any aqueous solution at 25°C , no matter what it contains, the product of $[\text{H}^+]$ and $[\text{OH}^-]$ must always equal 1.0×10^{-14} .

There are three possible situations.

- A) A neutral solution, where $[\text{H}^+] = [\text{OH}^-]$.
- B) An acidic solution, where $[\text{H}^+] > [\text{OH}^-]$.
- C) A basic solution, where $[\text{OH}^-] > [\text{H}^+]$.

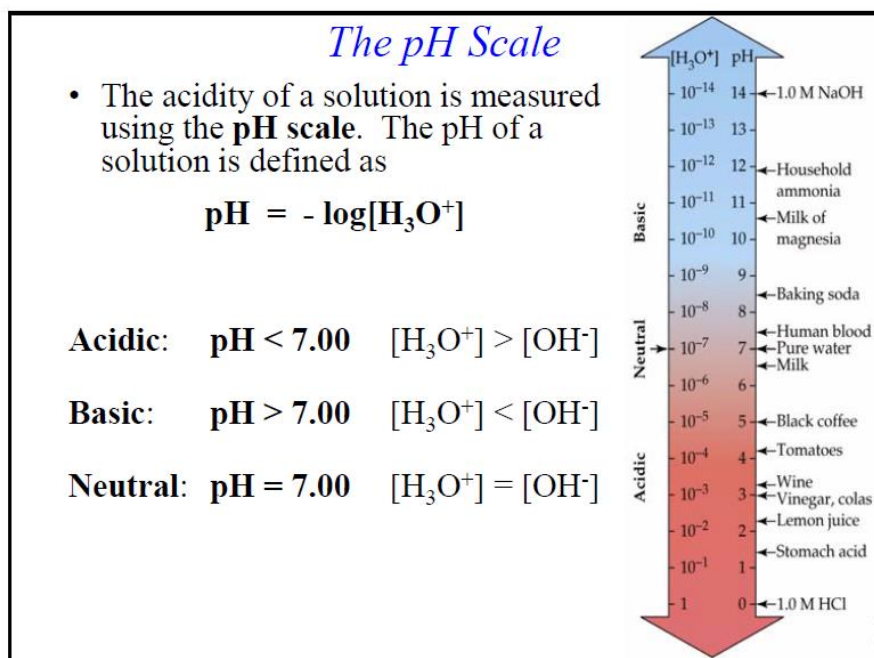
8.4 The pH scales

The concentration of H^+ in aqueous solution is usually very small and inconvenient to express and to draw as graphs. Sørensen, in 1909, developed the pH scale to solve this problem. The pH of a solution is defined as the negative logarithm of the hydrogen ion concentration (in mol/L):

$$\text{pH} = -\log [\text{H}_3\text{O}^+] \text{ or } \text{pH} = -\log [\text{H}^+]$$

Note that the negative logarithm gives us positive numbers for pH. pH decreases as the concentration of H^+ ions increase; in other words, the more acidic the solution, the lower its pH; the more basic the solution, the higher its pH.

The pH notation has been extended to other exponential quantities. For example, hydroxide concentrations can be expressed in terms of pOH, where $\text{pOH} = -\log [\text{OH}^-]$. Similarly, pKw can be expressed as $-\log K_w$.

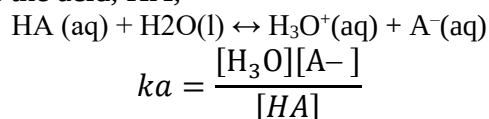


8.5 Measures of the Strengths of Acids and Bases in Aqueous Solution

1. Concentration of hydrogen and hydroxide ions
2. pH and pOH
3. Percent Ionization
4. Dissociation (Ionization) Constants

8.5.1 Acid Dissociation Constant, K_a

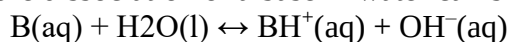
The acid-dissociation constant is a quantitative measure of the strength of the acid in a given solvent. For the dissociation of the acid, HA;



The K_a value for a strong acid is very large, because $[\text{HA}]$ in the denominator is very small since the acid dissociates essentially completely. Similarly, the K_a value for a weak acid is relatively small, since $[\text{H}_3\text{O}^+]$ and $[\text{A}^-]$ are very small, and $[\text{HA}]$ in the denominator is large.

8.5.2 Base dissociation constant, K_b

In the same way as for acids, the dissociation of a base in water can be written as



which gives the equilibrium expression of the form:

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

where K_b is the base dissociation constant. K_b values for strong bases are large, while K_b values for weak bases are small.

8.6 The Common ion Effect

The common-ion effect is caused by the addition of a compound having an ion in common with the dissolved substance that shift the equilibrium. The presence of the common ion suppresses the ionization of a weak acid or a weak base.

8.7 Buffer Solutions

A buffer solution is a solution that resists changes in pH from the addition of a limited amount of an acid or a base.

Buffers contain either a weak acid and its conjugate base or a weak base and its conjugate acid.

A buffer solution must contain a relatively large concentration of acid to react with any OH^- ions that are added to it. Similarly, it must contain a relatively large concentration of base to react with any H^+ ions.

Buffers are very important to chemical and biological systems. The pH in the human body varies greatly from one fluid to another. For example; the pH of blood is about 7.4, whereas the gastric juice

in our stomach has a pH of about 1.5. The pH values, which are crucial for proper enzyme function and the balance of osmotic pressure are maintained by buffers in most cases.

8.8 HYDROLYSIS OF SALTS

The word “hydrolysis” is derived from the Greek words hydro, meaning “water”, and lyses, meaning “to split apart”. Solutions of salts contain two types of ions, namely Cations and anions. The term salt hydrolysis describes the reaction of an anion or a cation of a salt, or both, with water. Depending on the strengths of the parent acids and bases, the cation of a salt can serve as an acid, base or neutral.

- ✓ Hydrolysis of Salts of Strong Acids and Strong Bases form neutral salt solution.
- ✓ Hydrolysis of Salts of Weak Acids and Strong Bases form basic salt solution.
- ✓ Hydrolysis of Salts of Strong Acids and Weak Bases form acidic salt solution.
- ✓ Hydrolysis of Salts of Weak Acids and Weak Bases: The solution of these salts contains both a strong acid (the cation of the weak base) and a strong base (the anion of the weak acid). Whether the solution of such a salt is acidic, basic or neutral depends on the relative strengths of the acidic cation and the basic anion. If the acid is stronger than the base, the solution is acidic and if the base is stronger than the acid, the solution is basic. If they are of equal strengths, the solution is neutral.

8.9 Acid–Base Indicators

Acid–Base Indicators are weak organic acids or weak organic bases that indicate whether a solution is acidic, basic, or neutral.

Indicators are used in the laboratory for estimating the pH of a solution and to reveal the equivalence point of a titration. Equivalence point is the point at which stoichiometrically equivalent quantities of an acid and a base have been brought together.

8.10 Equivalents of Acids and Bases

According to the definition of normality, the number of equivalents is the normality multiplied by the volume of solution, in liters. If we add enough acid to neutralize a given volume of base, the following equation holds:

$$N_1V_1 = N_2V_2$$

Where N_1 and V_1 refer to the normality, and volume of the acid solution, respectively, and N_2 and V_2 refer to the normality and volume of the base solution, respectively.

8.11 Acid–Base Titrations

A titration is a technique in which a solution of known concentration is used to determine the concentration of an unknown solution.

Titrant (the *known solution*) is added from a burette to a known quantity of the **analyte** (the *unknown solution*) until the neutralization reaction is complete.

The point at which the acid has completely reacted with or been neutralized by the base, or vice versa, is called **the equivalence point of the titration**.

Knowing the volume of **titrant** added allows the determination of the concentration of the unknown. Often, an indicator is used to signal the end of the reaction, the **end point**. The end point of titration is the pH at which the indicator changes color. A graph of pH as a function of the added titrant is called a **titration curve**.

8.11.1 Strong Acid-Strong Base Titrations

Features of the titration curve for the titration of a strong acid with a strong base.

- ✓ The pH is low at the beginning of the titration.
- ✓ The pH change slowly until just before the equivalence point.
- ✓ Just before the equivalence point, the pH rises sharply.
- ✓ At the equivalence point, the pH is 7.00.
- ✓ Just past the equivalence point, the pH continues its sharp rise.
- ✓ Further beyond the equivalence point, the pH continues to increase, but much less slowly.
- ✓ Any indicator whose colour changes in the pH range from about 4 to 10 can be used in the titration of a strong acid with a strong base.

Methyl violet changes colour too soon, and alizarin yellow R too late. So, bromothymol blue and phenolphthalein are preferred for this titration.

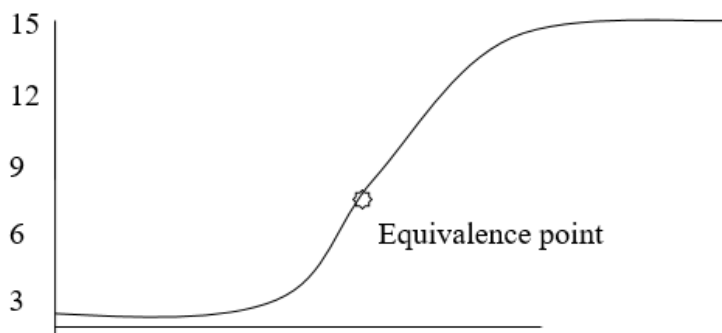


Fig. titration curve for the titration of a strong acid with a strong base

8.11.2 Weak Acid-Strong Base Titrations

The titration of a weak acid by a strong base is slightly more complicated than the titration of a strong acid by a strong base. The conjugate base of a weak acid will undergo hydrolysis, which will affect the pH of the solution. Thus, we need to consider the stoichiometric reaction between the acid and the base and the equilibrium reaction of the species that remain.

In contrast to the titration of a strong acid with a strong base, the titration of a weak acid with a strong base, has these features:

- ✓ The initial pH is higher because the weak acid is only partially ionized.
- ✓ At the half-neutralization, $\text{pH} = \text{pK}_a$. The solution at this point is a buffer solution in which the concentration of the weak acid and its conjugate base are equal.
- ✓ The pH is greater than 7 at the equivalence point because the anion of the weak acid hydrolyzes.

- ✓ The steep portions of the titration curve just prior to and just beyond the equivalence point is confined to a smaller pH range.
- ✓ The choice of indicator for the titration is more limited. The color change must occur in a basic solution. Generally, the midpoint of the pH range in which the indicator changes colour must be well above pH 7.

8.12 Questions

- The K_b of ammonia solution is 1.8×10^{-5} . What is the percent ionization of 0.1M CH_3COOH
 - 1.35 %
 - 4.23%
 - 0.42 %
 - 0.35%
- Which of the following is **not** amphoteric species?
 - H_2O
 - H_2SO_4
 - HCO_3^-
 - H_2PO_4^-
- Which of the following is the conjugate base of H_2S ?
 - HS^-
 - S^{2-}
 - SO_3^{2-}
 - H_3S
- Which of the following acid is the strongest?
 - HCN , $K_a = 7.2 \times 10^{-10}$
 - CH_3COOH , $K_a = 1.8 \times 10^{-5}$
 - HCOOH , $K_a = 1.8 \times 10^{-4}$
 - $\text{C}_6\text{H}_5\text{COOH}$, $K_a = 6.3 \times 10^{-3}$
- A buffer solution cannot be prepared from a mixture of: -
 - HCOOH and HCOOK
 - CH_3COOH and CH_3COOK
 - NH_3 and NH_4Cl
 - KOH and K_2SO_4
- All of the following salts from basic solution except
 - NaCO_3
 - CH_3COONa
 - NaCN
 - NaNO_3
- For the reaction $\text{HCl} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$, the Bronsted - Lowry acids are
 - H_3O^+ and HCl
 - HCl and H_3O^+
 - HCl and Cl^-
 - H_2O and Cl^-
- Which of the following is correct about a solution having a pH of 3?
 - the solution is basic
 - the solution has $[\text{H}^+] = 1 \times 10^3$
 - the solution has $[\text{OH}^-] = 1 \times 10^{-11}$
 - pH value of the solution is half of pK_w
- In the strong acid and strong base titration, The pH value at equivalent point is:-
 - less than 7
 - equal to 7
 - equal to 14
 - greater than 7
- When sodium acetate is added to an aqueous solution of acetic acid. Which of the Following is not true about the property of the resulted solution
 - The value of pH is increased
 - The Hydrogen ion concentration increased
 - The acetate ion suppress the ionization of acetic acid
 - Acetate ion concentration increased
- Which of the Following anions is the weakest Bronsted – Lowry base?
 - CO_3^{2-}
 - NO_3^-
 - PO_4^{3-}
 - F^-
- Which of the Following solution has the highest concentration of hydronium ions?

- A. $P^H=2$ B. $P^{OH}= 11$ C. $[H^+] = 0.001$ D. $[OH^-] = 0.001$
13. In a titration of weak base, the standard solution should be
 A. Strong Acid B. Strong Base C. Weak Base D. Indicator
14. The Strong acid has low
 A. $[H^+]$ C. Percent ionization
 B. PH value D. Dissociation constant
15. In titration of acids and bases, the point at which stoichiometrically equal amounts of substances have been brought together is known as
 A. End point B. Equivalence point C. Standard point D. Isomeric point.
16. Which of the Following substance undergoes hydrolysis in aqueous solution?
 A. NH_4OH B. H_2SO_4 C. HCl D. NH_4Cl
17. How is a buffer solution prepared?
 A. By mixing a weak acid and its conjugate base.
 B. By mixing a weak acid and strong acid.
 C. By mixing a strong acid and its conjugate base.
 D. By mixing a strong base and its conjugate acid.
18. To stabilize a solution's P^H add a(an)
 A. Buffer B. Weak acid C. Indicators D. Standard solution
19. The compound whose water solution has the highest P^H is
 A. $NaNO_3$ B. NH_4Cl C. $(NH_4)_2 SO_4$ D. Na_2CO_3
20. According to Lewis an acid is
 A. Proton acceptor C. Electron pair acceptor
 B. Proton donors D. Electron pair donor
21. Below the following table, which solutions are alkaline?

1	2	3	4	5	6	7	8	9	10	11	12	13	14
				↑		↑		↑			↑		
				A		B		C			D		

- A. C and D B. A and D C. A and B D. B and C

8.13 ANSWERS

1	A	2	B	3	A	4	D	5	D	6	D	7	B
8	C	9	B	10	C	11	B	12	A	13	A	14	B

15	B	16	D	17	A	18	A	19	D	20	C	21	A
----	---	----	---	----	---	----	---	----	---	----	---	----	---

9 UNIT 3 INTRODUCTION TO THERMODYNAMICS

Unit Outcomes

After completing this unit, you will be able to:

- ✓ understand the terms system, open system, closed system and isolated system, spontaneous process, state and path function, intensive and extensive property, internal energy, heat, and work;
- ✓ understand the first and second laws of thermodynamics;
- ✓ compare the energy change observed when chemical bonds are formed and broken, and relate these changes to endothermic and exothermic reactions;
- ✓ identify ways in which the terms reactant, product, and heat are combined to form thermodynamic equations representing endothermic and exothermic chemical change;
- ✓ understand concepts such as enthalpy change, entropy change and free-energy changes

9.1 INTRODUCTION TO THERMODYNAMICS

Thermodynamics is the study of energy and its transformation in a chemical reaction. Thermodynamics determines whether a reaction is possible and how much product can be formed. Thermodynamics deals with energy changes and applies to physical as well as chemical changes. However; it does not give any information about the duration of time required for the related processes. A thermodynamic system may be a solution in a beaker, a gas in a cylinder, a reaction vessel, an engine, an electrochemical cell, a biological cell, etc.

There are three types of thermodynamic systems based on the boundary between the system and its surroundings.

→ **An open system** – is a system that can exchange mass and energy, usually in the form of heat, with its surroundings. For example, an open system may consist of a quantity of water in an open container.

→ **A closed system** – is a system which transfers energy, but not mass across its boundary with its surroundings.

→ **An isolated system** – is a system in which neither the transfer of mass nor that of energy takes place across its boundary with the surroundings.

A system is said to be in a certain state if T , P , V , n (observable variables) do not change with time. An equation that relates these observable variables is called **an equation of state**.

An ideal-gas equation is an example of equation of state. A thermodynamic function (expressed in quantity) that depends on the initial and final state of a system, regardless of how that condition was achieved, is called a **state function**. Examples are energy, volume, Pressure, etc. Functions that depend on the path are **path functions**. Examples are heat, work, etc.

The properties of a system can be **intensive or extensive** based on its dependence on amount of a substance.

Intensive properties are those that do not depend on the amount of matter present. Some examples include color, temperature, and density.

Extensive properties are those properties that depend on the amount of matter, for example, mass, internal energy, volume, and pressure.

A **spontaneous** process is a process that occurs without external influence. Once started, requires no action from outside in order to continue.

On the other hand, a **non-spontaneous** process will not occur unless some external action is continuously applied.

An **isothermal** process is a process in which temperature remains constant. This typically occurs when a system in contact with an outside thermal reservoir (heat bath), and the change occurs slowly enough to allow the system to continually adjust to the temperature of the reservoir through heat exchange

9.2 THE FIRST LAW OF THERMODYNAMICS AND SOME THERMODYNAMIC QUANTITIES

Internal Energy (E): -of a system is the total energy contained within a system, partly as kinetic energy, and partly as potential energy.

When a chemical system changes from reactants to products and the product return to the starting reactant, the internal energy has changed.

To determine this change, ΔE , we measure the difference between the systems internal energy after the change (E_f) and before the change (E_o). $\Delta E = E_f - E_o$

The kinetic energy component of internal energy consists of various types of molecular motion and the movement of electrons within molecules.

Potential energy is a factor of:

- ✓ Attractive and repulsive interaction between molecules;
- ✓ Attractive interactions between electrons and nuclei within molecules;
- ✓ Repulsive interactions between electrons within molecules; and
- ✓ Repulsive interactions between nuclei within molecule.

9.2.1 Heat (Q)

Heat (thermal energy) is the energy transferred between a system and its surroundings as a result of a difference in their temperatures.

Heat passes spontaneously from the region of higher temperature to the region of lower temperature. Heat transfer stops when the system and surroundings reach the same temperature and therefore the system and surroundings are at **thermal equilibrium**.

Heat and Calorimetry

A chemical reaction or a physical change that releases heat is said to be an **exothermic** process. A reaction or a physical change that absorbs heat is an **endothermic** process. The SI unit of heat is joule (J). Another unit of heat is calorie.

A **calorie** (Cal) is the amount of heat, or other energy, necessary to raise the temperature of 1 g of water by 1 degree Celsius. A kilocalorie (kCal) is 1000 calories, and 4.184 Joules is equivalent to 1.0 calorie. The measurement of heat flow is called **Calorimetry**.

An apparatus that measures heat is called a **calorimeter**. The temperature changes experienced by an object when it absorbs a certain amount of energy is determined by its heat capacity.

Heat capacity is defined as the amount of heat energy required to raise its temperature by 1° (or 1 K). For pure substances, heat capacity is usually given for a specified amount of the substance.

The heat capacity of 1 mol of a substance is called its **molar heat** capacity. In all thermodynamic calculations temperature must be expressed in Kelvin (K). The heat capacity of 1g of a substance is called its **specific heat**.

$$\text{Specific heat} = q/m \Delta T$$

9.2.2 Work (W)

Work, like heat, is an energy transferred between a system and its surroundings.

There are several types of work, but here we will consider only the pressure-volume work that is done when gases are compressed or expanded.

Work is defined as a force applied over a given distance, so if the piston moves a distance Δh , then the magnitude of the work is $\text{Work} = \text{force} \times \text{distance} = F \times \Delta h$. Since $P = FA$, or $F = P \times A$, then $\text{Work} = F \times \Delta h = P \times A \times \Delta h$.

The product of pressure and a change in volume is usually called “pressure-volume” work. When a gas expands, ΔV is positive, and the work is negative.

Work (W) = $-P\Delta V$. A **negative** quantity of work signifies that the system loses energy, and that this, energy is transferred from the system to the surroundings.

When a gas is compressed by its surroundings, ΔV is negative, the quantity of work is positive and energy is gained by the system.

The First Law of Thermodynamics states that the total amount of energy in the universe is constant. When a system undergoes any chemical or physical change, the accompanying change in its internal energy, ΔE , is given by the heat added to the system, q , plus the work done by the system.

$$\Delta E = q + W$$

Thermodynamic quantities always consist of two parts:

- ✓ A number plus units, which gives the magnitude of the change; and
- ✓ A sign that indicates the direction of the flow.

The sign reflects the system's point of view. For example, if a quantity of energy flows into the system as heat, q is equal to $+q$, where the positive sign indicates that the system's energy is increasing. On the other hand, when energy flows out of the system as heat, q equal to $-q$, where the negative sign indicates that the system's energy is decreasing. This suggests that we need sign conventions to be used with the first law of thermodynamics.

The conventions are as follows:

- ✓ If heat is absorbed by the system, $q > 0$.
- ✓ If work is done on a system, $W > 0$.
- ✓ If heat is given off by a system, $q < 0$. If work is done by a system, $W < 0$.

Thermochemistry is the study of energy changes that occur during chemical reactions. Usually, we assess these energy changes in terms of losing and gaining energy in a system. When we warm our cold hands over a campfire, the burning wood gives off energy as heat and our hands gain energy, raising their temperature.

9.2.3 Heats of Reactions

In the first law of thermodynamics, you have learned that, $\Delta E = q + p\Delta V$. If a reaction occurs at constant volume, then $\Delta V = 0$ and no pressure-volume work is done.

What is enthalpy?

Enthalpy is defined as the sum of the internal energy and the pressure-volume product of a system.

$$H = E + PV \text{ or } \Delta H = \Delta E + P\Delta V \text{ (at constant pressure)}$$

$$\Delta E = q - P\Delta V \text{ then } \Delta H = q - P\Delta V + P\Delta V$$

$\Delta H = (q)_P$ Thus, the change in enthalpy equals the heat gained or lost at constant pressure

$$\Delta H = \Delta E + \Delta nRT \text{ or } \Delta E = \Delta H - \Delta nRT$$

9.2.4 Standard States

The standard state for a substance is the most stable form at temperature of 25°C and at atmospheric pressure of 1 atm.

The standard enthalpy of formation, $^\circ \Delta H$, is the change in enthalpy in the reaction when one mole of a substance is formed from the elements in their standard states. By definition, the standard enthalpy of formation of the most stable form of any element is zero.

Hess's Law states that if different processes are used to bring about the same change the enthalpy changes during these processes are the same.

Bond Energies

Bond energy is the enthalpy required to break a particular bond in 1 mole of gaseous molecules. This energy is generally expressed in kilojoules per mole of bonds (kJ/mol).

It is easy to understand bond-dissociation energy (D) for diatomic molecules because there is only one bond (single, double, or triple) per molecule.. Notice that the bond energy is always a positive quantity; energy is always required to break chemical bonds. Conversely, energy is released when a bond forms between two gaseous atoms or molecular fragments. Of course, the greater the bond energy, the stronger the bond.

$$\Delta H = \sum (\text{bond energies of bonds broken}) - \sum (\text{bond energies of bonds formed})$$

9.2.5 Entropy and Spontaneous Process

A spontaneous process is a change in a system that proceeds **without any outside influence** on the system. For example, when we add solid salt to water, the solid dissolves without any outside influence. We know well that certain other processes do not occur by themselves. They are said to be **non-spontaneous**. For example, water does not freeze to form ice at room temperature.

Entropy and Entropy Change

Entropy (S) is the thermodynamic property of a system that is related to its degree of **randomness**, or **disorder**. The greater the randomness, or disorder, in a system, the higher is its entropy. The SI unit of entropy is Joule per Kelvin, which is the same unit as that of heat capacity.

Entropy of a system affected by: -

- ✓ temperature

- ✓ state of substance
- ✓ hardness
- ✓ complexity of a substance

The entropy change, ΔS , for a chemical change is equal to the sum of the entropies of the products minus the sum of the entropies of the reactants.

$$\Delta S_{\text{rxn}} = \sum mS(\text{product}) - \sum nS(\text{reactant})$$

where m and n are the amounts of the individual species represented by their coefficients in the balanced equation.

- ✓ $\Delta S > 0$ indicates an increase in randomness, or disorder, during the change.
- ✓ $\Delta S < 0$ indicates a decrease in randomness, or an increase in order.

9.2.6 The Second Law of Thermodynamics

States that the entropy of the universe increases in a spontaneous process and remains unchanged in equilibrium process. Because the universe is made up of the system and its surroundings, the entropy change in the universe (ΔS_{univ}) for any process is the sum of the entropy changes in the system (ΔS_{sys}) and the surroundings (ΔS_{sur}).

Free Energy and Free Energy Change

The Gibbs free energy, G, is defined as

$G = H - TS$ where T is the absolute temperature, S is an entropy and H is an enthalpy.

For a process occurring at constant temperature, the change in free energy is given by the expression

$$\Delta G = \Delta H - T\Delta S.$$

- ✓ If $\Delta H < 0$ and $\Delta S > 0$ the change in Gibbs free energy becomes less than zero ($\Delta G < 0$), and the process is spontaneous.
- ✓ If $\Delta H > 0$ and $\Delta S < 0$, the change in Gibbs free energy becomes positive ($\Delta G > 0$), and the process is non-spontaneous.
- ✓ If $\Delta H > 0$ and $\Delta S > 0$, ΔG may or may not be negative, but only if $\Delta H < T\Delta S$, $\Delta G < 0$.
- ✓ If $\Delta H < 0$ and $\Delta S < 0$, $\Delta G < 0$ if $T\Delta S$ is small. This occurs at small value of temperature.
- ✓ If ΔG is zero, the process is at equilibrium and there is no driving force tending to make the reaction go in either direction.

9.3 Question

1. There is energy transfer but no matter exchange. What type of system is involved?
A. Open system B. Isolated system C. Closed system D. Isochoric system
2. Which of the following is spontaneous process?
A. Dissolving salt in water C. Resulting of iron in moist air
B. Flow of heat from hot body to cold body D. All are correct.
3. At what temperature is required 1.88KJ heat is absorbed by 15g of water from initially 20 °C at constant pressure? ($C_{\text{water}} = 4.18\text{J/g} \cdot ^\circ\text{C}$)
A. 90°C B. 50°C C. 70°C D. 30°C
4. Which of the following conditions lead to increase in the internal energy of a system?
A. 4J of heat are gained by the system and 8J of work are done by the system.
B. 8J of heat are gained by the system and 4J of work are done by the system.
C. 10J of heat are released by the system and 8J of work are done on the system.
D. 6J of heat are released by the system and 2J of work are done on the system.
5. How much work. In KJ is done when a gas compressed from 10L to 25L under a constant of 4 atm?
A. -60KJ B. -6.08KJ C. 6.08KJ D. 60KJ
6. At constant T and P, which of the following statements about the enthalpy(ΔH) and internal energy (ΔE) changes is connect when 2 moles of CO is converted to 2 moles of CO_2 for the reaction. $2\text{CO(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{CO}_2\text{(g)}$ ($R = 8.314\text{J/mol.K}$)
A. $\Delta H \neq \Delta E$ B. $\Delta H = \Delta E$ C. $\Delta H < \Delta E$ D. $\Delta E > \Delta H$
7. Which of the following process is the sign of Entropy change is negative?
A. Sublimation of iodine
B. Formation of silver chloride precipitate from a solution of sodium chloride and sodium nitrate
C. Evaporation of sea water
D. Dissolving of potassium chloride
8. Based on the following equation ΔH , ΔS and ΔG results spontaneous toward low temperature. Which one?
A. $2\text{O}_3\text{(g)} \rightarrow 3\text{O}_2\text{(g)}$ C. $2\text{C(graphite)} + 2\text{H}_2\text{(g)} \rightarrow \text{C}_2\text{H}_4\text{(g)}$
B. $2\text{H}_2\text{O(g)} \rightarrow 2\text{H}_2\text{(g)} + \text{O}_2\text{(g)}$ D. $\text{N}_2\text{(g)} + 3\text{H}_2\text{(g)} \rightarrow 2\text{NH}_3\text{(g)}$
9. In chemical equilibrium, what will happen when products favored over reactants?

- A. $\Delta G^\circ > 0$ and $K < 1$ B. $\Delta G^\circ < 0$ and $K > 1$ C. $\Delta G^\circ > 0$ and $K > 0$ D. $\Delta G^\circ < 0$ and $K < 1$
10. Which of the following is not a state function?
- A. Volume B. Internal energy C. Free energy D. Work
11. In a certain chemical process both ΔH and ΔS have values greater than zero. Under what conditions, the reaction would not be spontaneous?
- A. $\Delta H > T\Delta S$ B. $T\Delta S - \Delta H > 0$ C. $\Delta H = T\Delta S$ D. None
12. Which of the following is true for a change in Gibbs free energy?
- A. For $\Delta H < 0$ and $\Delta S > 0$, the $\Delta G < 0$, and the process is spontaneous.
 B. if $\Delta H > 0$ and $\Delta S < 0$, $\Delta G > 0$ and the process is non-spontaneous.
 C. If ΔG is zero, the process is at equilibrium
 D. ALL
13. A system that can exchange mass and energy, usually in the form of heat, with its surroundings is
- A. An open system B. closed system C. An isolated system D. none
14. In thermodynamics, a quantity whose value simply depends upon the initial and final state of the system is called:
- A. thermodynamic quantity C. state function
 B. adiabatic quantity D. path function
15. Which of the following is true?
- A. If work is done by a system, $W < 0$ C. If heat is given off by a system, $q < 0$.
 B. If work is done on a system, $W > 0$. D. all
16. Which of the following leads to decrease in the entropy?
- A. $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$ C. $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
 B. $\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s}) \rightarrow \text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{aq})$ D. all
17. What is the enthalpy of the formation of gaseous hydrazine in the rxn
- $$\text{N}_2(\text{g}) + 2\text{H}_2(\text{g}) \rightarrow \text{N}_2\text{H}_4(\text{g})?$$
- (Bond Energy of N-N=163kJ, H-H=436kJ, N-H=389kJ and N≡N=945kJ)
- A. 99kJ B. 1933kJ C. 1818kJ D. None
18. What is the enthalpy for the following reaction? (Note reverse equation (2) and add to equation (1)) $\text{C}(\text{graphite}) \rightarrow \text{C}(\text{diamond})$
- Given the data: 1. $\text{C}(\text{graphite}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}); \Delta H = -393.5 \text{ kJ}$
 2. $\text{C}(\text{diamond}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}); \Delta H = -395.4 \text{ kJ}$

A. 1.9KJ B. 19KJ C. 788.9KJ D. None
 19. What is ΔH_f^0 for the rxn $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$? (ΔH_f^0 of $\text{CH}_4 = -84.75\text{KJ/mol}$, $\text{O}_2 = 0\text{KJ/mol}$, $\text{CO}_2 = -393.8\text{KJ/mol}$ and $\text{H}_2\text{O} = -285.8\text{KJ/mol}$)

A. -890.55 KJ/mol B. -904.8KJ/mol C. -965.4 KJ/mol D. None
 20. Which of the f/g is not an intensive physical property of substances?

A. color B. temperature C. density D. volume
 21. Which of the f/g is non- spontaneous process?

- | | |
|---|--|
| A. Melting of ice at 25°C | C. reverting iron oxide to iron and oxygen |
| B. reaction of NaCl with HCl | D. adding solid salt to water |

9.4 Answer

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

10 UNIT 4 ELECTRO CHEMISTRY FOR GRADE 12th

Unit Outcomes

After completing this unit, you will be able to:

- ✓ understand the fundamental concepts related to oxidation-reduction reactions;
- ✓ know the application of redox reactions in the production of new substances and electrical energy;
- ✓ demonstrate an understanding of fundamental concepts related to the inter conversion of chemical and electrical energy;
- ✓ understand the difference between metallic conduction and electrolytic conduction;
- ✓ identify and describe the function of the components of Electrolytic and Galvanic cells;
- ✓ understand the difference between electrolytic and Galvanic cells;
- ✓ know how to solve problems based on Faraday's first and second laws;
- ✓ measure, through experimentation, the mass of the metal deposited by electroplating (example copper to copper (II) sulphate), and apply Faraday's Law to relate the mass of the metal deposited to the amount of charge passed;
- ✓ predict spontaneity of redox reactions and overall cell potentials by studying a table of half-cell reduction potentials;
- ✓ determine the emf of an electrochemical cell, experimentally or from given data;
- ✓ explain the application of electrochemistry in our daily lives and in industry;
- ✓ explain corrosion as an electrochemical process, and describe corrosion-inhibiting techniques (e.g., painting, galvanizing, cathodic protection)

10.1 ELECTROCHEMISTRY

Electrochemistry is the study of production of electricity from the energy released during a spontaneous chemical reaction and the use of electrical energy to bring about non-spontaneous chemical transformations

10.2 Electrochemical cells

A spontaneous chemical process is the one which can take place on its own and in such a process the Gibb's energy of the system decreases. It is this energy that gets converted to electrical energy. The reverse process is also possible in which we can make non-spontaneous processes occur by supplying external energy in the form of electrical energy. These inter conversions are carried out in equipment's called Electrochemical Cells.

10.3 Types of electrochemical cells

Electrochemical Cells are of two types:

1. **Galvanic Cells;** Converts chemical energy into electrical energy
2. **Electrolytic Cells;** Converts electrical energy into chemical energy

10.3.1 Electrolytic cell

These electrodes are dipped in an electrolytic solution containing cations and anions. On supplying current, the ions move towards electrodes of opposite polar.

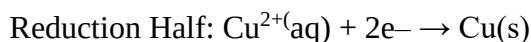
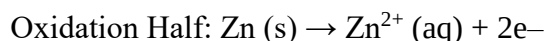
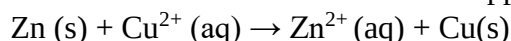
Preferential Discharge of ions

Where there are more than one cation or anion the process of discharge becomes competitive in nature. Discharge of any ion requires energy and in case of several ions being present the discharge of that ion will take place first which requires the energy

10.3.2 Galvanic cell

Cell energy is extracted from a spontaneous chemical process or reaction and it is converted to electric current.

For example, Daniel Cell is a Galvanic Cell in which Zinc and Copper are used for the redox reaction to take place.



Zn is the reducing agent and Cu^{2+} is the oxidizing agent. The half cells are also known as *Electrodes*.

The oxidation half is known as Anode and the reduction half is called Cathode. Electrons flow from anode to cathode in the external circuit. Anode is assigned negative polarity and cathode is assigned positive polarity. In Daniel Cell, Zn acts as the anode and Cu acts as the cathode.

10.4 Electrode potential

It may be defined as the tendency of an element, when it is placed in contact with its own ions to either loss or gain electrons and in turn become positively or negatively charged. The electrode potential will be named as oxidation or reduction potential depending upon whether oxidation or reduction has taken place

- ✓ Both oxidation and reduction potentials are equal in magnitude but opposite in sign.
- ✓ It is not a thermodynamic property, so values of E are not additive

10.5 Standard electrode potential (E^0)

It may be defined as the electrode potential of an electrode determined relative to standard hydrogen electrode under standard conditions.

The standard conditions taken are:

- ✓ 1M concentration of each ion in the solution.
- ✓ A temperature of 298 K
- ✓ 1 bar pressure for each gas

10.6 Electrochemical series

The half-cell potential values are standard values and are represented as the standard reduction potential values as shown in the table at the end which is also called Electrochemical Series

10.7 Cell potential or emf of a cell

The difference between the electrode potentials of two half cells is called cell potential. It is known as electromotive force (EMF) of the cell if no current is drawn from the cell.

$$E_{\text{cell}} = E_{\text{cathode}} + E_{\text{anode}}$$

For this equation we take oxidation potential of anode and reduction potential of cathode. Since anode is put on left and cathode on right, it follows

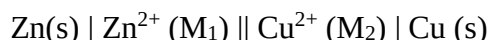
$$\text{Therefore, } E_{\text{cell}} = E_R + E_L$$

For a Daniel cell, therefore

$$E_{\text{cell}}^0 = E_{\text{Cu}^{2+}/\text{Cu}} - E_{\text{Zn}/\text{Zn}^{2+}}^0 = 0.34 + (0.76)$$

10.8 Cell Diagram or Representation of a Cell

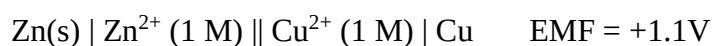
The following conventions or notations are applied for writing the cell diagram in accordance with IUPAC recommendations. The Daniel cell is represented as follow



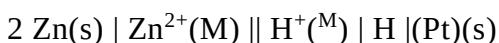
- ✓ Anode half-cell is written on the left-hand side while cathode half-cell on right hand side.
- ✓ A single vertical line separates the metal from aqueous solution of its own ion



- ✓ A double vertical line represents salt bridge
- ✓ The molar concentration is placed in brackets after the formula of the corresponding ion.
- ✓ The value of e.m.f. of the cell is written on the extreme right of the cell. For example

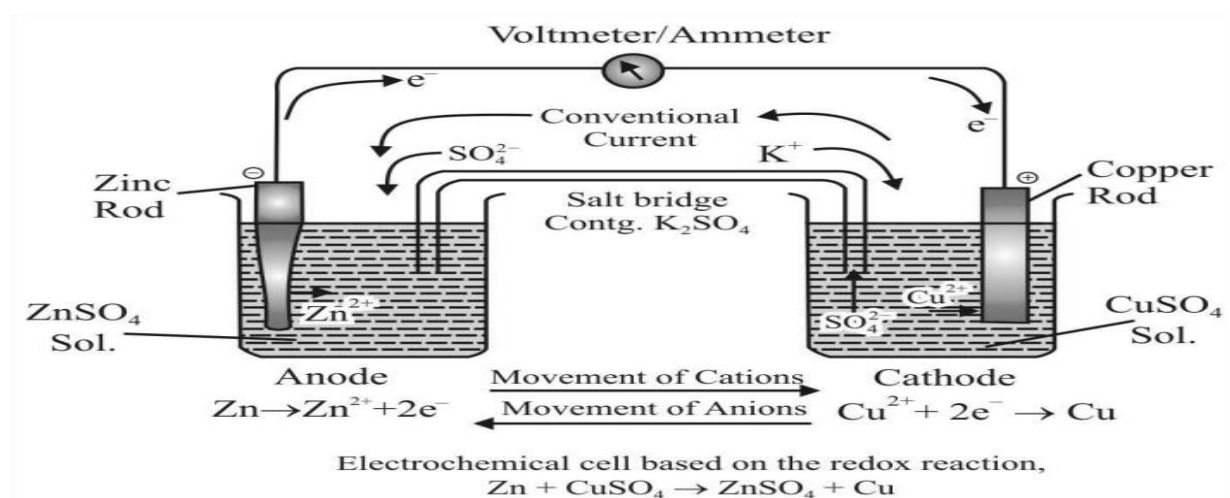


- ✓ If an inert electrode like platinum is involved in the construction of the cell, it may be written along with the working electrode in bracket say for example, when a zinc anode is connected to a hydrogen electrode.



10.9 4.8. Salt bridge

Salt bridge is used to maintain the charge balance and to complete the circuit by facilitating the flow of ions through it. It contains a gel in which an inert electrolyte like Na_2SO_4 or KNO_3 etc are mixed. Negative ions flow to the anode and positive ions flow to the cathode through the salt bridge and charge balance is maintained and cell keeps on functioning



10.10 Spontaneity of a reaction

$$\Delta G = -nFE_{\text{CELL}}$$

For a spontaneous cell reaction ΔG should be negative and cell potential should be positive.

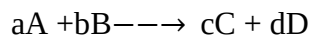
If we take standard value of cell potential in the above equation, we will obtain standard value of ΔG as well.

$$\Delta G^0 = -nFE^0_{\text{CELL}}$$

10.11 NERNST EQUATION

It relates electrode potential with the concentration of ions.

Thus, the reduction potential increases with the increase in the concentration of ions. For a general electrochemical reaction of the type.



Nernst equation can be given as $E_{\text{cell}} = E_{\text{cell}}^0 - \frac{RT}{nF} \ln \frac{[C]^c [D]^d}{[A]^a [B]^b}$

$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{2.303}{nF} RT \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Substituting the values of R and F we get

$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{0.0591}{n} \log \frac{[C]^c [D]^d}{[A]^a [B]^b} \text{ at } 298\text{K}$$

Equilibrium Constant from Nernst Equation

For a Daniel cell, at equilibrium

$$E_{\text{cell}} = 0 = E_{\text{cell}}^0 - \frac{2.303RT}{2F} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

$$E_{\text{cell}}^0 - \frac{2.303RT}{2F} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

$$\text{At equilibrium, } \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = K_C$$

$$E_{\text{cell}}^0 = \frac{2.303RT}{2F} \log K_C$$

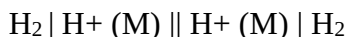
$$= \frac{2.303 \times 8.314 \times 298}{2 \times 96500} \log K_C$$

$$\text{In general, } E_{\text{cell}}^0 = \frac{0.0591}{n} \log K_C$$

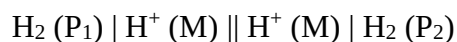
10.12 Concentration cells

If two electrodes of the same metal are dipped separately into two solutions of the same electrolyte having different concentrations and the solutions are connected through salt bridge, such cells are known as concentration cells.

For example



✓ Electrode concentration cells



$$E^0_{\text{cell}} = \frac{0.0591}{n} \log \frac{P_2}{P_1}, \text{ where } P_2 < P_1 \text{ for spontaneous reaction}$$

✓ **Electrolyte concentration cell**

The EMF of concentration cell at 298 K is given by

$$\text{Zn} | \text{Zn}^{2+} (\text{M}_1) || \text{Zn}^{2+} (\text{M}_2) | \text{Zn}$$

$$E^0_{\text{cell}} = \frac{0.0591}{n} \log \frac{M_2}{M_1}, \text{ where } M_2 > M_1 \text{ for spontaneous reaction}$$

10.13 BATTERIES

When Galvanic cells are connected in series to obtain a higher voltage the arrangement is called Battery.

Primary Batteries

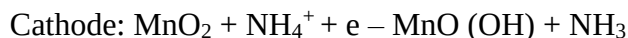
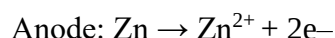
Primary cells are those which can be used so long the active materials are present. Once they get consumed the cell will stop functioning and cannot be re-used. Example Dry Cell or Leclanché cell and Mercury cell

Dry cell

Anode: Zn container

Cathode: Carbon (graphite) rod surrounded by powdered MnO_2 and carbon.

Electrolyte: NH_4Cl and ZnCl_2 Reaction:



The standard potential of this cell is 1.5 V and it falls as the cell gets discharged continuously and once used it cannot be recharged

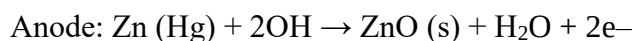
Mercury cells

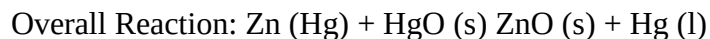
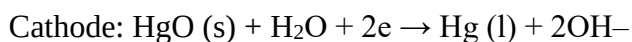
These are used in small equipment's like watches, hearing aids.

Anode: Zn – Hg Amalgam

Cathode: Paste of HgO and carbon

Electrolyte: Paste of KOH and ZnO





The cell potential is approximately 1.35V and remains constant during its life.

Secondary Batteries

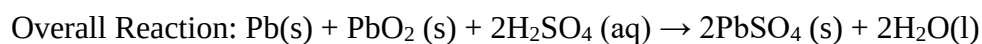
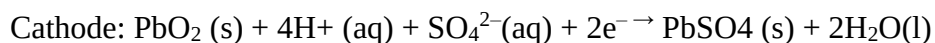
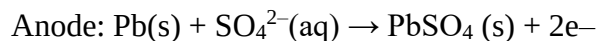
Secondary cells are those which can be recharged again and again for multiple uses. e.g. lead storage battery and Ni – Cd battery

Lead Storage Battery

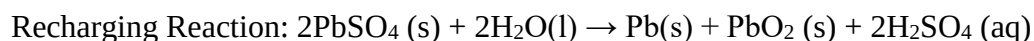
Anode: Lead (Pb)

Cathode: Grid of lead packed with lead oxide (PbO₂)

Electrolyte: 38% solution of H₂SO₄ Discharging Reactions

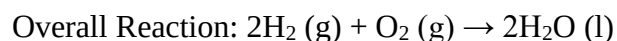
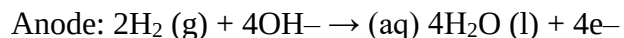
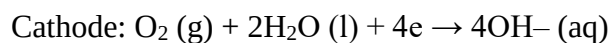


To recharge the cell, it is connected with a cell of higher potential and this cell behaves as an electrolytic cell and the reactions are reversed. Pb(s) and PbO₂(s) are regenerated at the respective electrodes. These cells deliver an almost consistent voltage.



Fuel cells

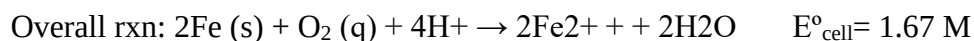
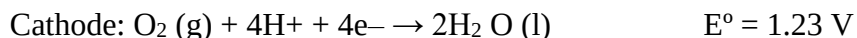
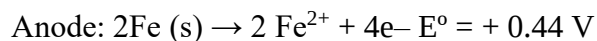
A fuel cell differs from an ordinary battery in the sense that the reactants are not contained inside the cell but are externally supplied from an external reservoir. Fuel cell is used in space vehicles and in this cell the two gases are supplied from external storages. In this cell carbon rods are used as electrodes with KOH as the electrolyte.



10.14 CORROSION

It involves a redox reaction and formation of an electrochemical cell on the surface of iron or any other metal. At one location oxidation of iron takes place (anode) and at another location

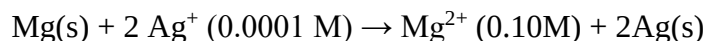
reduction of oxygen to form water takes place (cathode). First Fe gets oxidized to Fe^{2+} and then in the presence of oxygen it forms Fe^{3+} which then reacts with water to form rust which is represented by $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$



Rusting of iron can be avoided by painting it or by coating it with some other metals like Zinc. The latter process is known as Galvanization. As the tendency of Zn to get oxidized is more than iron it gets oxidized in preference

10.15 Questions

- The two advantages of $\text{H}_2\text{—O}_2$ fuel cell over ordinary cell.
 - They do not cause any pollution
 - They have high efficiency
 - A&B
 - None
- The process of coating zinc over iron is called _____.
 - Galvanizations
 - Electroplating
 - Cathodic protection
 - Painting
- A metal which is more electropositive than iron such as Al, Zn, Mg can be used in cathode protection of iron against rusting. This process is known as:-
 - Galvanizations
 - Electroplating
 - Cathodic protection
 - Painting
- The following chemical reaction is occurring in an electrochemical cell.



The E° electrode values are

$$\text{Mg}^{2+} / \text{Mg} = -2.36 \text{ V}$$

$$\text{Ag}^+ / \text{Ag} = 0.81 \text{ V}$$

For this cell

- Cell potential E_{cell} is 2.96V
- Symbolic representation of the above cell is $\text{Mg} | \text{Mg}^{2+} (0.1 \text{ M}) || \text{Ag}^+ (0.0001\text{M}) |$
- All are true statement

- A cell is prepared by dipping copper rod in 1 M copper Sulphate solution and zinc rod in 1 M zinc sulphate solution. The standard reduction potentials of copper and zinc are 0.34 V and -0.76 V respectively. What will be the standard electromotive force (EMF) of the cell?
A. 0.76V B. 0.34V C. 1.10V D. -1.10V
- The standard free energy change and maximum work obtainable for the reaction occurring in the cell: Daniel cell. $\text{Zn (s)} \mid \text{Zn}^{2+} (1 \text{ M}) \parallel \text{Cu}^{2+} (1 \text{ M}) \mid \text{Cu (s)}$. Given $E^\circ_{\text{Zn}^{2+} / \text{Zn}} = -0.76\text{V}$, $E^\circ_{\text{Cu}^{2+} / \text{Cu}} = 0.34 \text{ V}$, $F 96,500 \text{ C mol}^{-1}$. Also calculate the equilibrium constant for the reaction
A. $2.21 \times 10^5 \text{ J/mol}$ B. $-2.21 \times 10^5 \text{ J/mol}$ C. $2.21 \times 10^{-5} \text{ J/mol}$ D. $2.21 \times 10^{-5} \text{ J/mol}$
- In an electrolysis experiment, a current was passed for 5 hours through two cells connected in series. The first cell contains a solution gold salt and the second cell contains copper sulphate solution. 9.85 g of gold was deposited in the first cell. If the oxidation number of gold is +3, find the amount of copper deposited on the cathode in the second cell. Also calculate the magnitude of the current in ampere
A. 0.0804 ampere B. - 0.0804 ampere C. 0.0404 ampere D. -0.0404 ampere
- Which of the following has been universally accepted as a reference electrode at all temperatures and has been assigned a value of zero volt?
A. platinum electrode C. reduction takes place at the anode
B. copper electrode D. standard hydrogen electrode
- Which of the following statement is wrong about galvanic cell?
A. cathode is positive charged C. reduction takes place at the anode
B. anode is negatively charged D. reduction takes place at the cathode
- The standard electrode potentials (reduction) of Pt/Fe^{3+} , Fe^{2+} and Pt/Sn^{4+} , Sn^{2+} are + 0.77 V and 0.15 V respectively at 25°C . The standard EMF of the reaction $\text{Sn}^{4+} + 2\text{Fe}^{2+} \rightarrow \text{Sn}^{2+} + 2\text{Fe}^{3+}$ is
A. -0.62 V B. -0.92 V C. $+0.31 \text{ V}$ D. $+0.85 \text{ V}$
- Strongest reducing agent is:
A. K B. Mg C. Al D. I
- Zn cannot displace following ions from their aqueous solution:

- A. Ag^+ B. Cu^{2+} C. Fe^{2+} D. Na
12. A metal having negative reduction potential when dipped in the solution of its own ions, has a tendency:
- A. to pass into the solution C. to become electrically positive
B. to be deposited from the solution D. to remain neutral None
13. E° for the half-cell reactions are as,
- A. $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$; $E^\circ = + 0.76 \text{ V}$
B. $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$; $E^\circ = + 0.41 \text{ V}$. The E° for half-cell reaction, $\text{Fe}^{2+} + \text{Zn} \rightarrow \text{Zn}^{2+} + \text{Fe}$ is
- A. $- 0.35 \text{ V}$ B. $+ 0.35 \text{ V}$ C. $+ 1.17 \text{ V}$ D. $- 0.17 \text{ V}$
14. The emf of the cell; $\text{Ti}/\text{Ti}^+ (0.0001 \text{ M}) \parallel \text{Cu}^{2+} (0.01 \text{ M})/\text{Cu}$ is 0.83 V . The emf of this cell will be increased by:
- A. Increase the concentration of Cu^{++} ions C. Increasing the concentration of both
B. Decreasing the concentration of Ti^+ D. A and B both
15. In electroplating the article to be electroplated is made
- A. cathode C. either cathode or anode
B. anode D. simply suspended in the electrolytic bath
16. On electrolyzing a solution of dilute H_2SO_4 between platinum electrodes, the gas evolved at the anode is
- A. SO_2 B. SO_3 C. O_2 D. H_2 .
17. A spoon to be electroplated with gold should be:
- A. cathode B. anode C. electrolyte D. none of these
18. Three faradays of electricity was passed through an aqueous solution of iron (II) bromide. The mass of iron metal (at. mass 56) deposited at the cathode is –
- A. 56 g B. 84 g C. 112 g D. 168 g
19. The electric charge for electrode deposition of one gram equivalent of a substance is :
- A. one amp/sec B. 96,500 C/sec C. one amp/hour D. 96,500 C
20. Number of electrons involved in the electrodeposition of 63.5 g of Cu from a solution of CuSO_4 is :

- A. 6.022×10^{23} B. 3.011×10^{23} C. 12.044×10^{23} D. 6.022×10^{22}
21. When one coulomb of electricity is passed through an electrolytic solution the mass deposited on the electrode is equal to :
- A. equivalent weight B. molecular weight C. electrochemical equivalent D. one gram
22. 13.5 g of Al get deposited when electricity is passed through the solution of AlCl_3 . The number of faradays used are :
- A. 0.50 B. 1.00 C. 1.50 D. 2.00
23. An ion is reduced to the element when it absorbs 6×10^{20} electrons. The number of equivalents of the ion is
- A. 0.50 B. 1.00 C. 1.50 D. 2.00
24. Electrolysis can be used to determine atomic masses. A current of 0.550 A deposits 0.55 g of a certain metal in 100 minutes. Calculate the atomic mass of the metal if $n = 3$:
- A. 100 B. 45.0 C. 48.25 D. 44.75
25. When a lead storage battery is discharged
- A. PbSO_4 is formed B. Pb is formed C. SO_2 is consumed D. H_2SO_4 is formed
26. As lead storage battery is charged
- A. lead dioxide dissolves C. sulphuric acid is regenerated
B. lead electrode becomes coated with lead sulphate D. the concentration of sulphuric acid
27. Which of the following displacement does not occur:
- A. $\text{Zn} + 2\text{H}^+ \rightarrow \text{Zn}^{2+} + \text{H}_2$ C. $\text{Cu} + \text{Fe}^{2+} \rightarrow \text{Cu}^{2+} + \text{Fe}$
B. $\text{Fe} + 2\text{Ag}^+ \rightarrow \text{Fe}^{2+} + 2\text{Ag}$ D. $\text{Zn} + \text{Pb}^{2+} \rightarrow \text{Zn}^{2+} + \text{Pb}$
28. The standard free energy change for the following reaction is -210 kJ . What is the standard cell potential? $2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$
- A. $+0.752$ B. $+1.09$ C. $+0.420$ D. $+0.640$
29. How many electrons are delivered at the cathode during electrolysis by a current of 1A in 60 seconds?
- A. 3.74×10^{20} B. 6.0×10^{23} C. 7.48×10^{21} D. 6.0×10^2
31. During the electrolysis of fused NaCl , the reaction that occurs at the anode is:
- A. Chloride ions are oxidized C. Chloride ions are reduced
B. Sodium ions are oxidized D. Sodium ions are reduced

10.16 Answers

1	C	2	A	3	C	4	D	5	C	6	B	7	B
8	D	9	C	10	A	11	A	12	D	13	A	14	B
15	D	16	A	17	C	18	A	19	B	20	D	21	C
22	C	23	C	24	C	25	C	26	A	27	B		

11 UNIT 5 CHEMISTRY OF SOME ELEMENTS AND COMPOUNDS FOR GRADE 10th&12th

Unit Outcomes

After completing this unit, you will be able to:

- ✓ describe the occurrence and abundance of the elements in nature;
- ✓ explain how carbon, nitrogen, and phosphorus cycles in nature;
- ✓ understand metallurgical processes;
- ✓ understand the occurrence, extraction, chemical properties and uses of sodium, calcium, tin, lead, zinc and chromium;
- ✓ describe the occurrence, extraction, chemical properties, and main uses of silicon;
- ✓ explain the major steps in the industrial production of ammonia, nitric acid, sulphuric acid and diammoniummonohydrogen phosphate.

11.1 Chemistry of Some Elements and Compounds

11.2 Metallurgy

- It is a science of extracting metals from their natural sources
 - Steps
 1. Concentrating ore :- is increasing the amount of metal containing done by
 - a. Gravity separation :- based on different in density
 - b. Magnetic separation :- concentrate ferromagnetic ores
 - c. Froth flotation:- concentrate sulphide ores
 - d. Amalgamation :- concentrate mercury ores
 - e. Leaching :- concentrate Al and Au ores
 2. Pretreatment :- is converting an ore to form easy to reduce
 3. Production of metals :- metals are manufacture by reduction process
 - Chemical reduction: - using reducing agents
e.g., Less active metals such as Cu, Zn, Cr, etc.
 - Electrolytic reduction using electrolysis
e.g., active metals such as alkaline earth metals, Al, etc.
 4. Refining and alloying: - by electrolysis, distillation and zone refining
 - * Roasting: - is heating ore in the presence of air to oxide
 - * Calcination: - is heating an ore in the absence of air

11.3 Important metals

No	Metals	Occurrence	Extraction	Chemical property	Use
----	--------	------------	------------	-------------------	-----

1	Sodium (Na)	-In combined state -Main ore, NaCl -Other are: borax, trona, salt pitter	Downs process	-React with -water, air, halogens and hydrogen	-Cool fuels -Lighting high ways
2	Calcium (ca)	-main ore CaCl_2 -other minerals are: lomestone, dolomite, gypsum	Electrolysis of molten CaCl_2	- react with water, air, halogens	- in heart actions -as reducing agent -remove gases(O_2, N_2) from electron is tube
3	Tin (Sn)	Cassiterite, SnO_2	Chemical reduction with carbon	-react with air, steam, HCl , con. HNO_3 , Con. NaOH , Cl_2 -amphoteric metals	-to make tin plate -to make solder (Sn + Pb) -bronze (Sn+Cu)
4	Lead (Pb)	Main ore is galena, PbS	-roasting of PbS to PbO and reducing coke -self reduction	-react with air, Cl_2 and S - amphoteric metals	- to make electrode for lead storage batteries -to make lead sheets, pipes, and solder
5	Zink (Zn)	Main ore zinc blend (ZnS)	By chemical reduction	-furnish in air -React with acid and alkalis -show amphoteric	-To coat iron -to make alloys and electrodes for dry cell
6	Chromium (Cr)	Main ore Chromite, FeCr_2O_4	By chemical reduction with coke	-react with air, HCl , and H_2SO_4	- to make nichrome wire -to make alloys such as stainless steel -to make chromic acid(H_2CrO_4)

11.4 Alloys

No	Alloy	Components
1	Ferrochrome	Fe and Cr

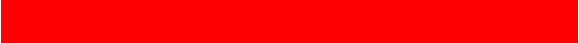
2	Bronze	Cu and Sn
3	Brass	Cu and Zn
4	Nichrome	Ni and Cr
5	Solder	Sn and Pb
6	Type metals	Sn,Pb and Sb
7	Stainless steel	Fe,Cr and Ni

11.5 Important compounds

No	compound	Properties	Use
Ammonia(NH ₃)	Haber process	-Colorless gas -irritating odour -extremely soluble in water -react with oxygen	-To make fertilizer -to make HNO ₃ , dyec, etc -as heat exchange fluid in large refrigeration units
Nitric acid	Ostwald process	-Is powerful oxidizing agent -colorless, oily liquid	-to make ammonium nitrate -to make nylon, explosives such as trinitrotoluene (TNT), nitrocellulose
Sulphuric acid	Contact process	-viscous liquid -powerful dehydrating agent -it a strong diprotic acid	- to make fertilizer, papers, explosives, etc.

11.6 Questions Grade 10th&12th (important inorganic metals and compounds)

1. Which metal is manufactured by chemical reduction?

- A. K B. Zn C. Al D. Na
2. Which metal is purified by distillation?
- A. Hg B. Cu C. Al D. Sn
3. Which metal is used to cool fuel in nuclear reactors?
- A. Ca B. Li C. Pb D. Na
4. All of the following are minerals of sodium EXCEPT?
- A. Borax B. trona C. dolomite D. Chile salt peter
5. What is the purpose of CaCl_2 in the production of sodium by down process?
- A. To lower the melting point C. 
- B. to higher melting point of NaCl D. 
6. Which of the following is used for galvanization of iron sheets?
- A. Na B. Zn C. Pb D. Si
7. Which of the following acid that dissolves silica?
- A. HF B. HNO_3 C. H_2SO_4 D. HCl
8. Which of the following is not amphoteric metals?
- A. Zn B. Sn C. Pb D. Cr
9. Which metal is manufactured by electrolytic reduction?
- A. Pb B. Sn C. Mg D. Cu
10. Which of the following alloy components is type metal?
- A. Sn ,Pb and Sb B. Cr, Hg, Au C. Ni and Cr D. Sn and Pb
11. Why ammonia is completely soluble in water, -----?
- A. due to high reactivity C. due to ionic bonding
- B. due to hydrogen bonding D. due to less reactivity
12. which of the following happen to the nitric oxide formed in the third step during Ostwald process?
- A. Reduction of ammonia C. recycled in the second step
- B. Oxidation of ammonia D. heating sodium nitrate
13. Which one of the following metals extracted from its ore by self-reduction?
- A. Pb B. Sn C. Cr D. Zn
14. Which one of the following methods is used to concentrate an ore that is denser than its gangue?

- A. Magnetic separation
 - B. gravity separation
 - C. froth flotation
 - D. leaching
28. What is the main use of adding Na_3AlF_6 in to Al_2O_3 in the metallurgy of aluminum?
- A. to concentration Al_2O_3
 - B. to decrease the melting point of Al_2O_3
 - C. to remove unnecessary impurities from Al_2O_3
 - D. all the above
29. the leading industrial chemical that is used as an electrolyte in car battery is
- A. HNO_3
 - B. HCl
 - C. H_2SO_4
 - D. H_3PO_4
30. Salt that absorb water from the atmosphere, but remain solid id known as
- A. Deliquescent
 - B. hygroscopic
 - C. efflorescent
 - D. all of the above
31. Which of the following methods is used to decrease environmental pollution?
- A. burning waste materials in open air away from a town
 - B. burring the waste in the space at high altitude
 - C. burning the waste deep in the ground
 - D. recycling industrial and agricultural waste
32. why silicate minerals not commonly used as a source of metals
- A. they are difficult to reduce and concentrate
 - B. they are available minerals
 - C. do not contain important metals
 - D. found in depth in the ocean
33. aluminum is used in electrical cables due to its
- A. resistance to corrosion
 - B. brightness and large abundance
 - C. evolution of heat when burned
 - D. Lightness and electrical conductivity
34. the main raw materials of cements are
- A. Soda ash & silica
 - B. soda ash & clay
 - C. lime stone & clay
 - D. lime stone& silica
35. The purpose of treating skin with brine solution during tanning is to
- A. Increase durability of the tanned skin
 - B. Remove hair from the skin
 - C. remove flesh from the skin

- ### 11.7 Answer to review question Grade 10th&12th on important metals and compounds

109

15	C	16	C	17	B	18	C	19	D	20	B	21	A
22	C	23	B	24	C	25	D	26	D	27	D	28	C
29	C	30	B	31	D	32	A	33	D	34	C	35	D
36	D	37	A	38	D	39	A	40	C	41	A	42	C
43	D												

12 UNIT 6 POLYMER CHEMISTRY FOR GRADE 12th

Unit outcomes

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

- ✓ define polymer;
- ✓ define, identify, and classify monomers;
- ✓ list a variety of synthetic & natural polymers & explain their differences;
- ✓ classify synthetic polymers as addition or condensation polymers & identify their monomers;
- ✓ describe the process of addition & condensation polymerization;
- ✓ describe the monomers, properties & uses of plastics, rubbers, carbohydrates & proteins

12.1 6. POLYMER

A polymer is a large molecule or a macromolecule which essentially is a combination of many subunits (monomers). Polymers are not restricted to monomers of the same chemical composition or molecular weight and structure. Some natural polymers are composed of one kind of monomer. Most natural and synthetic polymers, however, are made up of two or more different types of monomers; such polymers are known as copolymers.

Most synthetic polymers are organic compounds. Examples are nylon, Dacron, Teflon, Bakelite, polyethylene and pvc. **A polymer** is a large molecule or a macromolecule which essentially is a combination of many subunits (monomers). Polymers are not restricted to monomers of the same chemical composition or molecular weight and structure. Some natural polymers are composed of one kind of monomer. Most natural and synthetic polymers, however, are made up of two or more different types of monomers; such polymers are known as **copolymers**.

Copolymers can be random or regular, based on the way the monomers are arranged. Polymers can be found all around us. From the strand of our DNA which is a naturally occurring biopolymer to polypropylene which is used throughout the world as plastic. Most synthetic polymers are organic compounds. Examples are nylon, Dacron, Teflon, Bakelite, polyethylene and pvc.

Monomer	Polymer	Polymer Name	Some Uses
$\text{CH}_2=\text{CH}_2$	$\sim\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\sim$	polyethylene	plastic bags, bottles, toys, electrical insulation
$\text{CH}_2=\text{CHCH}_3$	$\begin{array}{ccccccc} \sim\text{CH}_2 & \text{CH} & \text{CH}_2 & \text{CH} & \text{CH}_2 & \text{CH} & \sim \\ & & & & & & \\ & \text{CH}_3 & & \text{CH}_3 & & \text{CH}_3 & \end{array}$	polypropylene	carpeting, bottles, luggage, exercise clothing
$\text{CH}_2=\text{CHCl}$	$\begin{array}{ccccccc} \sim\text{CH}_2 & \text{CH} & \text{CH}_2 & \text{CH} & \text{CH}_2 & \text{CH} & \sim \\ & & & & & & \\ & \text{Cl} & & \text{Cl} & & \text{Cl} & \end{array}$	polyvinyl chloride	bags for intravenous solutions, pipes, tubing, floor coverings
$\text{CF}_2=\text{CF}_2$	$\sim\text{CF}_2\text{CF}_2\text{CF}_2\text{CF}_2\text{CF}_2\text{CF}_2\sim$	polytetrafluoroethylene	nonstick coatings, electrical insulation

Polymerization can be classified as:

Addition (chain growth) polymerization made by addition reaction like polyethylene, polystyrene, pvc,

Condensation (step growth) polymerization made by reaction of monomers involved two or more functional groups Example polyester, nylons, Bakelite.

Based on their response to heat polymers can be

Thermoplastics (can be soften on heating and moulded in to d/t shapes) and

Thermosetting plastics (once solidified and set they cannot be softened or remolded by heating).

One of the features that we value in plastics, their durability, is also a problem for us. When plastics are disposed into the environment, they remain there for a very long time, degrading very slowly, and continue to pollute our environment.

The production of plastic produces substantial amounts of toxic chemicals (e.g. ethylene oxide, benzene and xylenes) to air and water. Many of the toxic chemicals released in plastic production can cause cancer and birth defects and damage the nervous system, blood, kidneys and immune systems. These chemicals can also cause serious damage to environment.

12.2 Question

1. A polymer which is used for making hoses for chemicals, patroles, and oils b/s of its resistance to chemicals is:
A. Neoprene B. SBR C. PTFE D. Bakelite
2. The monomer(s) of **Bakelite** is/ are
A. Adipic acid and hex methylenediamine C. Phenol and formaldehyde
B. Propylene and phenol D. Styrene and phenol
3. All of the following are examples of **condensation** polymers except
A. PVC B. PET or Dacron C. Nylon D. Bakelite
4. Which of the following is a **thermosetting** plastic?
A. Bakelite B. Teflon C. Polyester D. PMM
5. The widely used polymer used for making packing materials and preparing CD DVD toys flowerpots is
A. Polystyrene B. Polyvinyl chloride C. Perspex D. PTFE
6. The polymer used in food preparation by coating cooking utensils is
A. Polystyrene B. polytetrafluoroethylene (Teflon) C. PMMA D. Dacron
7. Which of the following is the first **synthetic fiber**?
A. Dacron B. Rayon C. Bakelite D. Teflon
8. PVC has ---**CH₂CHClCH₂CHClCH₂CHCl**---structure (At.wt.H=1, Cl=35.5, C=12). How many **monomer** units are present in PVC that has a molar mass of 1.33×10^5 gm/mol?
A. 2128 B. 1228 C. 709 D. 2228
9. The monomer of **natural rubber** is
A. styrene B. Isoprene C. Chloroprene D. Isobutene
10. A polymer used to produce making squeeze bottles, plastic wrapping, and garment bags is
A. Polyethylene B. Polypropylene C. PVC D. Rayon
11. Which of the f/g is produced by the condensation reaction of **diacids** and **diamines**?
A. Nylons B. Teflon C. Rayon D. Dacron
12. Which of the following is an animal starch?

- A. Maltose B. Cellulose C. Glucose D. Glycogen

13. Which of the following bimolecular forms at higher or lower pH?

- A. Carbohydrate B. Protein C. Fat and oils D. Polymer

12.3 Answers for unit 6 polymer

1 A	2 B	3 A	4 A	5 A	6 B	7 B
8 A	9 B	10 A	11 A	12 D	13 B	

13 Grade 10 SOME IMPORTANT IN ORGANIC COMPOUNDS

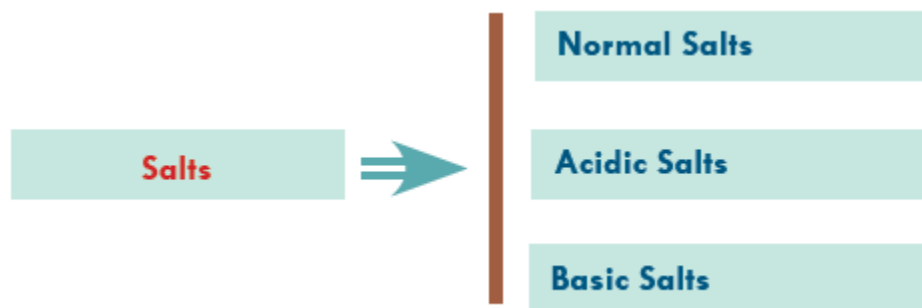
Unit outcomes

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

- ✓ Classify inorganic compounds on the basis of their composition and/or their chemistry;
- ✓ Discuss the types of oxides and their chemical properties ;
- ✓ Mention the classification of salts;
- ✓ Explain plant nutrients & their compositions.

13.1 SOME IMPORTANT INORGANIC COMPOUNDS

- ✓ **Acidic oxides** are, therefore, known as acid anhydrides, e.g., sulfur dioxide is sulfurous anhydride; sulfur trioxide is sulfuric anhydride. When these oxides combine with bases, they produce salts.
- ✓ Generally Group 1 and Group 2 elements form bases called base anhydrides or basic oxides.
- ✓ **Basic oxides** are the oxides of metals. If soluble in water they react with water to produce hydroxides (alkalis). E.g. Na_2O , K_2O , CaO , MgO
- ✓ An amphoteric solution is a substance that can chemically react as either acid or base.
- ✓ **Amphoteric oxides** are metallic oxides, which show both basic as well as acidic properties. E.g. Al_2O_3 , ZnO , PbO
- ✓ These are the oxides, which show neither basic nor acidic properties, that is, they do not form salts when reacted with acids or bases, e.g., carbon monoxide (CO); nitrous oxide (N_2O); nitric oxide (NO), etc., are neutral oxides.
- ✓ **Oxides:** Group 1 metals react rapidly with oxygen to produce several different ionic oxides, usually in the form of M_2O . With the oxygen exhibiting an oxidation number of -2.
- ✓ **Peroxides:** Often Lithium and Sodium reacts with **excess** oxygen to produce the peroxide, M_2O_2 . with the oxidation number of the oxygen equal to -1.
- ✓ **Superoxide's:** Often Potassium, Rubidium, and Cesium react with excess oxygen to produce the superoxide, MO_2 , with the oxidation number of the oxygen equal to -1/2.
- ✓ Basic oxide, acidic oxide and amphoteric oxides are salt forming oxides.
- ✓ Monoprotic, diprotic and polyprotic acids are classified depending on number of replaceable hydrogen ions per molecules.



Salts	Uses
Sodium chloride (common salt, NaCl)	1. Preparation and preservation of food 2. Raw material for the manufacture of sodium chloride, and sodium hydroxide 3. Component of oral rehydration salt (ORS) 4. Manufacture of baking soda (NaHCO_3) AND Na_2CO_3 5. In making a freezing mixture which is used by ice cream vendors
Ammonium nitrate, (NH_4NO_3)	Nitrogenous fertilizer and explosives
Copper (II) sulphate, (CuSO_4)	1. Used to make Bordeaux mixture (mixture of CuSO_4 and Ca(OH)_2 and other fungicides. Bordeaux mixture is used to prevent fungal attack of leaves and vines 2. Useful in electroplating and dyeing
Iron (III) chloride (FeCl_3)	1. Used in the treatment of waste water 2. Used for etching printed circuits
Potassium nitrate, (KNO_3)	1. Used in making gun powder ((mixture of KNO_3 carbon and sulphur) and other explosives 2. Used as fertilizer in agriculture
Calcium sulphate, ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$, Gypsum)	Use for plastering walls and supporting fractured bones
Barium sulphate, (BaSO_4)	1. Given to patients as a barium meal before gastrointestinal x-ray photography 2. Used as a white pigment
Iron (II) sulphate (FeSO_4)	Given as iron tablets to patients who suffer from anemia

Properties of salts

Salts can be classified depending on the anion (negative ion) they possess, because the anion is partly responsible for the solubility the salt.

1. Solubility of salts

When salts are added to water or water is added to salt, different types of actions take place. Some salts are soluble in water; some are insoluble while some salts are only slightly soluble.

2. Tendency to absorb water from the atmosphere or release water to the atmosphere salts can be classified as hygroscopic, deliquescent and efflorescent depending on their tendency to absorb water from or release water to the atmosphere

Hygroscopic salts are those which absorb water from the atmosphere but remain solid.

Example: anhydrous copper (II) sulphate, CuSO_4

Deliquescent salts absorb water from the atmosphere to form a solution. The process of absorbing water from the atmosphere by a solid to form a solution is called deliquescence.

Example: calcium chloride (CaCl_2), sodium nitrate (NaNO_3), iron (III) chloride (FeCl_3)

Efflorescent salts lose their of crystallization to the atmosphere. The loss of water of crystallization by solid crystals to the atmosphere is known as efflorescence. It is very important to note that all deliquescent substances are hygroscopic, but all hygroscopic substances are not necessarily deliquescent.

Macronutrients	micronutrients
These are the elements required in relatively large amounts by plants. These elements are nitrogen, Phosphorus, potassium, calcium, magnesium, sulphur, carbon, hydrogen and oxygen.	These are the elements requires by plants in relatively small (trace) amounts. These elements are iron, cobalt, zinc, manganese, molybdenum, copper, boron and chlorine. The elements carbon, hydrogen and oxygen are obtained by plants from air and water. These are not considered as mineral nutrients.

Nitrogen (N)

- a. Is a very important plant nutrient and absorbed by plants in the form of nitrate ions (NO_3^-)
- b. Is utilized in the synthesis of amino acids, proteins, coenzymes and nucleic acids which are taking place in the growth and development of plants
- c. Is also involved in the synthesis of chlorophyll to produce deep green color.

Phosphorus (p)

- a. Is absorbed by plants mainly in the form of H_2PO_4^- and HPO_3^{2-} in small amounts.
- b. Facilitates early growth and root formation, quick maturity and promotes seed or fruit production.
- c. Has a role in the formation of some amino acids and proteins, coenzymes, nucleic acids and high energy phosphate compounds like Adenosine Triphosphate (ATP)

Potassium (K)

- a. Is absorbed by plants in the form of the K^+ ion
- b. Is a component of enzymes that facilitate the process of photosynthesis, protein synthesis, and adjusts stomata movement.
- c. Is the most important ion controlling the turgidity of plant cells.
- d. Increases in the solute potential (sap concentration) within the cells which leads to an increase in the amount of water that enters the cells osmotically.

Fertilizers are materials that are added to soils to increase the growth, yield or nutritional value of crops. There are organic (natural) and synthetic (artificial) fertilizers.

Organic (natural) fertilizers are fertilizers derived from animals and plants.

- ✓ Organic fertilizers provide readily available nutrients to plants after some period of decay and decomposition and can supply the elements nitrogen, phosphorus, potassium and various trace elements to the soil.
- ✓ Synthetic (artificial) fertilizers contain nitrogen, phosphorous and potassium, and are known as complete fertilizers.
- ✓ A mixed artificial fertilizer may contain nitrogen, phosphorus and potassium. Such type of fertilizer is called a complete fertilizer.
- ✓ The term NPK is used to describe fertilizers containing the elements nitrogen, phosphorus and potassium.

Examples of artificial fertilizers are ammonium sulphate, $(\text{NH}_4)_2\text{SO}_4$, potassium nitrate (KNO_3), ammonium nitrate (NH_4NO_3), and diammonium hydrogen phosphate, DAP ($(\text{NH}_4)_2\text{HPO}_4$). The most important mineral nutrients that need to be added to the soil are usually nitrogen, phosphorus and potassium.

- ✓ All three elements are needed in large quantities. Commercial fertilizers are normally given a "grade", which reflects the percentages they contain of N, P and K by dry weight. For example, a NPK fertilizer may be described as 10; 10; 10 indicating that there is 10% of each element in the fertilizer.
- ✓ The suitable proportions are best determined in relation to the tested fertility of the soil and the requirements of the particular crop that is being grown on it. This means fertilizers of the same composition are not used for different purposes.
- ✓ In nitrogen-deficient soil, a fertilizer containing a higher percentage of nitrogen than phosphorus and potassium must be used. In addition to this, all fertilizers are not used for the treatment of all kinds of soil. For example, ammonium sulphate is more suitable for use in basic soils than in acidic or neutral soils.

Pesticides

- ✓ Pesticides are chemicals that can be applied to crops to kill pests that affect plant growth and development.

Herbicides

- ✓ Herbicides (weed killers) are chemicals used to manipulate or control undesirable plants. Some examples of herbicidal salts are: Common salt (NaCl), copper sulfate (CuSO_4), sodium arsenite (NaAsO_2), Ammonium sulfamate ($\text{NH}_4\text{SO}_3\text{NH}_2$), sodium chlorate (NaClO_3).

13.2 Question

1. An example of a base which is not an alkali is_____?
A. Caustic potash
B. copper hydroxide
C. Caustic soda
D. All of these
2. Oxides which dissolve in water to form alkaline solution are called_____.
A. Neutral Oxide B. Amphoteric Oxide C. Acidic Oxide D. Basic oxide.
3. An amphoteric oxide is one which?
A. Reacts with acid only to form salt and water
B. Reacts with base only to form salt and water
C. Reacts with acid as well as alkali to form salt and water
D. Reacts with acid as well as alkali to form salt and hydrogen
4. An Oxide in which the two oxygen atoms are linked by –O-O- link and with atoms of other elements is known as
A. An Oxide B. A base C. An amphoteric D. Peroxide.
5. A salt formed by replacement of hydroxyl radical of a diacidic or tri acid base with an acid is: -
A. Normal salt B. Basic salt C. Acidic salt D. Double salt
6. During the experimental work for chemical test, the colour of the solution Fe (III) salt is generally?
A. Reddish brown precipitate
B. White precipitate
C. Pale-green precipitates
D. Yellow precipitate
7. An example of a deliquescent salt is?
A. Anhydrous calcium chloride
B. Quick lime
C. Blue vitriol
D. washing soda
8. When Oxygen combines with another element to produce binary compound is:
A. Acids B. Bases C. Salts D. Oxides
9. Take about 20 ml of solution of potassium hydroxide in conical flask and add to it few drops of universal indicator solution. The nature of the solution will turn?
A. Red B. Deep blue C. Orange –red D. Pink

10. When NO is dissolved in water the nature of the solution is
- A. Acidic B. Neutral C. Basic D. Peroxide.
11. A salt which absorbs moisture from the air, but does not change its physical state are called :
- A. Deliquescent salt B. Ternary salt C. Efflorescent salts D. Hygroscopic salts
12. Which of the following sets of compounds are salts?
- A. MgCl_2 , Na_2CO_3 , CaSO_4 C. NaOH , $\text{Mg}(\text{OH})_2$, $\text{Fe}(\text{OH})_3$
 B. HBr , HNO_3 , H_2SO_4 D. MgO , Fe_2O_3 , NO
13. A certain base that was purchased from a super market ionizes almost completely in water. Then it is a:
- A. Strong base B. weak base C. dilute base D. concentrated base.
14. Which of the following chemical reaction represents an acid-base reaction?
- A. $\text{HBr} + \text{KOH} \rightleftharpoons \text{KBr} + \text{H}_2\text{O}$ C. $\text{ZnCl}_2 + \text{MgSO}_4 \rightleftharpoons \text{ZnSO}_4 + \text{MgCl}_2$
 B. $\text{Na}_2\text{CO}_3 + \text{CaCl}_2 \rightleftharpoons \text{CaCO}_3 + 2\text{NaCl}$ D. $\text{NH}_4\text{OH} + \text{KCl} \rightleftharpoons \text{KOH} + \text{NH}_4\text{Cl}$
15. Base on the following flame taste, which one is incorrectly matched?
- A. Sodium $\rightarrow$ yellow C. Barium $\rightarrow$ green
 B. Calcium $\rightarrow$ crimson D. potassium $\rightarrow$ violet
16. Which of the following is used as a raw material in the production of sulfuric acid?
- A. NO_2 B. CO_2 C. SO_3 D. V_2O_5
17. An example of a base which is caustic for an alkali is_____.
- A. Lime water B. Copper hydroxide C. Caustic soda D. None
18. A salt prepared by precipitate (double displacement of two soluble salt)
- A. NaNO_3 B. AgCl C. NaCl D. K_2SO_4
19. Take a 200 ml beaker and pour in it 100ml of distilled water. To the water add 10 ml of concentrated drops by drop and continuously stir. These forms
- A. Dilute sulphuric solution C. concentrated sulphuric solution
 B. Saturated sulphuric solution D. Super saturated sulphuric solution

20. When concentrated sulphuric acid poured down the side of the test tube, the acid sinks the bottom and a brown ring is formed where the two layers meet. The formula of the substance that forms the brown ring is $\text{FeSO}_4 \cdot \text{NO}$, then used to identify the presence of :-

- A. NO_2^- B. NO_3^- C. SO_4^{2-} D. SO_3^{2-}

21. Which of the following set of elements are primary mineral nutrients?

- A. Zn, B, Fe B. Cl, Cu, Mo C. Fe, Cl, Mn D. N, P, K

22. Of the following which one is an inorganic pesticide?

- A. CuSO_4 B. HCN C. Borax D. phosphine

23. It adjusts stomata movements, what type of mineral nutrients in the plant growth?

- A. Nitrogen B. Potassium C. Phosphorus D. Oxygen

24. Which mineral nutrients required in small amount by plant for the growth and development?

- A. Calcium B. carbon C. Boron D. Sulphur

13.3 Grade 10 SOME IMPORTANT IN ORGANIC COMPOUNDS

1	B	2	D	3	C	4	D	5	B	6	A	7	A	8	D
9	B	10	B	11	C	12	A	13	A	14	A	15	B	16	C
17	C	18	B	19	A	20	B	21	D	22	A	23	B	24	C

14 GRADE 9 THE PHYSICAL STATES OF MATTER

Unit outcomes

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

- ✓ Explain the kinetic theory of matter;
- ✓ Describe the gas laws;
- ✓ Describe the solid, liquid & gas state;

14.1 The Kinetic Theory of Matter

Generally, the kinetic theory of matter is based on the following three assumptions:

1. All matter is composed of particles which are in constant motion.
2. The particles possess kinetic energy and potential energy.
3. The difference between the three states of matter is due to their energy contents and the motion of the particles.

14.2 THE GASEOUS STATE

Assumptions of the kinetic molecular theory of gases

1. The particles are in a state of constant, continuous, rapid, random motion and, therefore, possess kinetic energy.
2. The volume of the particles is negligible compared to the total volume of the gas.
3. The attractive forces between the particles are negligible. There are no forces of attraction or repulsion between gas particles.
4. The average kinetic energy of gas particles depends on the temperature of the gas.

14.3 The Gas Laws

Boyle's Law. Boyle's law states that the volume of a fixed mass of gas is inversely proportional to the pressure at a constant temperature. The inverse relationship between pressure and volume is mathematically given as $V \propto 1/P$ (at constant T and n)

If P_1 and V_1 represent the initial conditions; and P_2 and V_2 represent the new or final

Conditions,

Boyle's law can be written as:

$$P_1V_1 = P_2V_2$$

Charles' law States that the volume of a fixed mass of gas at constant pressure varies directly with the Kelvin temperature.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Combined gas law.

- ✓ Is a combination of Boyle's and Charles' law
- ✓ Expresses the relationship between pressure, volume, and temperature of a fixed amount of gas.

Derivation of the combined gas law:

Boyle's law: $V \propto 1/P$

Charles' law: $V \propto T$

Then, $V \propto T/P$ (combined)

$V = kT/P$ (where k is a constant)

The combined gas law equation is given as follows

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Real gases only obey the ideal gas laws closely at high temperature and low pressure. Under these conditions, their particles are very far apart.

The ideal gas law is a combination of Boyle's law, Charles' law and Avogadro's law.

Boyle's law: $V \propto 1/P$ (at constant T and n)

Charles' law: $V \propto T$ (at constant P and n)

Avogadro's law $V \propto n$ (at constant P and T)

This relationship indicates how the volume of gas depends on pressure, temperature and number of moles.

Where R, is proportionality constant called the gas constant.

$$PV = nRT \quad \text{Where } R = 0.082 \text{ Latm/mol.k}$$

Graham's law of diffusion states that at constant temperature and pressure, the rate of diffusion of a gas, r, is inversely proportional to the square root of its density, d, or molar mass, M.

$$\frac{r_1}{r_2} = \frac{\sqrt{d_2}}{\sqrt{d_1}}; \quad \frac{r_1}{r_2} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

He found that gases having low densities diffuse faster than gases with high densities.

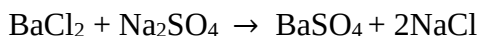
14.4 Energy Changes in Liquids

The rate of evaporation of a liquid depends on three factors: temperature, intermolecular forces and surface area

The temperature at which liquid and solid states coexist in equilibrium is called melting point of the solid or the freezing point of the liquid. Both take place at the same temperature.

14.5 Question

- Which one of the following is true about kinetic theory and properties of matter?
 - The particles are in a state of constant, continuous, rapid, random motion
 - The average kinetic energy of gas particles depends on the temperature of the gas.
 - The volume of the particles is negligible compared to the total volume of the gas
 - All
- Which of the following happens when a liquid freezes?
 - The particles will assume the shapes of the container
 - The particles form a more ordered arrangement
 - inter molecular forces decrease
 - intermolecular distance increase
- Rate of evaporation of a liquid depends on
 - intermolecular force
 - temperature
 - surface area
 - All
- Which one of the following is true?
 - boiling and evaporation are the same
 - if energy more than 40.6kJ is applied to water at 1.0atm, its boiling point will be less than 1000°C
 - liquids freezing of water and melting of ice take place at the same temperature
 - nonvolatile liquids have higher Vapor pressure than volatile
- Which of the following is true about the factors influencing rate of reaction?
 - As surface area of reactants decreases, the rate of a reaction decreases
 - As concentration of a reactant increases, the rate of the reaction increases except for zero-order reactions
 - As temperature increases, the rate of reaction decreases
 - A catalyst always speeds up the rate of a reaction
- What type of reaction does usually take place in the reaction of?



- A. Combination
B. single displacement
C. double displacement
D. single displacement
7. The Pre-conditions for a chemical reaction is
- A. Proper orientation B. Collision C. Activation energy D. All
8. Which one of the following is true about heat of reaction?
- A. For exothermic reactions, ΔH is negative because the energy of the reactants is greater than the energy of the products.
B. The enthalpy of the system decreases
C. For endothermic reactions, ΔH is positive because the energy of the product is higher than the energy of the reactant.
D. ALL
9. A sample of a gas occupies 312cm³ at 822kpa. what volume will the same gas occupies at 938 kpa, if the temperature is kept constant?
- A. 273.4cm³ B. 321Cm³ C. 283.4Cm³ D. 323cm³
10. The volume of a fixed mass of a gas increased from 250cm³ to 375cm³ when it was heated at a constant pressure. What was the original temperature if the final temperature was 1270C?
- A. 6.3 °c B. -127°c C. -6.3°c D. 127°c
11. The conditions on 166cm³ of a gas at 93000pa and 15⁰c are increased by 7kpa and 5⁰c. The new volume of the gas in liters is: -
- A. 167cm³ B. 157cm³ C. 288cm³ D. 323cm³
12. The pressure exerted by 48g of CH₄ in a 30L container at 27⁰C?
- A. 2.64atm B. 2.84atm C. 2.46atm D. 2.48atm
13. The relative rates of diffusion of oxygen and hydrogen gases?
- A. 4 B. 3 C. 2 D. 8

14.6 ANSWER GRADE 9 THE PHYSICAL STATES OF MATTER

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