COURSE CODE: CHEM 101

COURSE TITLE: General Chemistry I

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Chemistry Unit

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Course Description

- This course introduces students to the study of matter of which chemistry is a branch. It teaches the fundamental processes and phenomenon of chemistry such as: What is matter?, states of matter, measurements, concentration units, atomic structure, periodicity, gases, nomenclature, chemical bonding, colligative properties of matter , redox, acids/bases, chemical kinetics ,solutions and thermochemistry.

- The Prerequisite for this course includes: Secondary school chemistry, algebra or concurrent enrollment in General Mathematics 101.

News Forum

- YOU ARE EXPECTED TO CHECK THE NEWS FORUM FROM TIME TO TIME FOR ANY ANNOUNCEMENTS AND COURSE INFORMAITON PROVIDED BY THE TEACHER
- YOU MAY ALSO RESPOND TO SUCH INFORMATION OR REQUEST FROM THE TEACHER IN THE NEWS FORUM THROUGH THE 'QUESTIONS AND ANSWERS FORUM'.

Meet the teacher

- Hi Everyone,

- My name is Adebayo Omotayo Onigbinde

- I am an Associate Professor of Physical and Analytical Chemistry at the department of Basic Sciences, Babcock University, Ilishan, Remo, Ogun State

- I had my Bachelor's and Masters degree from Obafemi Awolowo University, Ile-Ife Oyo State, Nigeria and my Ph.D. from the University Of Delaware, Newark, DE 19716 in the United States of America.

- I teach general , physical, analytical and organic chemistry to the undergraduate science and medical students

- I also teach post graduate students from various departments in analytical chemistry courses.

- I had a lot of industrial and teaching experience in the United States of America and Nigeria and I joined Babcock University in August 2011.

- I am also a registered Nurse from the United States of America

Module Time Frame

- This course will be divided into modules with subunits
- You are advised to spend a minimum of one hour on each unit of the module
- The time frame to complete a module depend on the number of sub units within each module
- You are also expected to submit assignments given at the end of each module within two days of posting the assignment on the News Forum bulleting board
- Thanks and remain blessed

Module 1: Matter and Chemistry

In this module, we shall discuss the following topics:

- Bible and the study of Chemistry
- What is matter?
- States of matter
- What is chemistry?
- Branches of chemistry
- Chemistry and other disciplines
- What is a chemical?
- Chemical reactions
- Scientific methods
- Definition of technology

Module 1;Unit 1: cont'd

Creation Story: The Bible described the creation story in the book of Genesis. It says that on the first day God created Light by using "Words" and then separated the firmament into the Heavens and the Earth. Later, he created everything on earth including man and He rested (vacation) at the end of the exercise. (Genesis Chapter 1)

- The firmament, Earth, is what the scientist calls Matter
- Different approaches are used to study different aspect s of matter
- **What is matter?** – Everything on this earth that has mass and occupy space (volume) such as: air, water, plants, soil, forest, rays etc.
- **State of matter**: Matter exist in three states: gas, liquid, and solid
- **What is chemistry?**

    Chemistry is the study of matter in regards to it's:
    - Composition (what it is made of)
        - Structure (shape or form)
        - Properties (descriptive behavior)
    - Reactions (how it responds to different conditions)
- It is involved in everything we do in our daily activities (eating, washing clothes etc)

Module 1 Unit 1: cont'd

- **Branches of chemistry**:
- General/Analytical chemistry (composition/properties/reaction)
- Inorganic chemistry (compounds that do not contain Carbon)
- Organic chemistry (compounds containing Carbon)
- Physical chemistry (physical nature of matter)
- **Chemistry and other disciplines**
- Chemistry cuts across many disciplines of studies such as:
- Geology (geochemistry- soil, ores, minerals)
- Biology (biochemistry- chemistry of proteins)
- Agriculture (phytochemistry (antioxidants, proximate analysis)
- Pharmacy (pharmaceutical chemistry, pharmacokinetics) etc

Module 1 Unit 2
Chemicals, Chemical Reactions, Scientific Methods and Technology

What is a chemical? – Any substance with same composition/properties irrespective of where it comes from.

- **What is a chemical reaction?** Chemical reaction occurs when a chemical undergoes a change to form a new substance with new properties e.g.

$$\text{Gari + Hot water} \rightarrow \text{Eba}$$
$$2H_2 + O_2 \rightarrow 2H_2O$$
$$\text{Wood + fire} \rightarrow \text{Ashes}$$
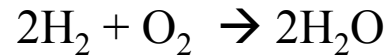
- Chemical reactions take place all the time and anywhere (library, kitchen, in the body (happy and bad mood), factories, and when you feel a sensation (falling in and out of love) etc.
- Scientist uses scientific methods to study science including chemistry.
- **What are Scientific Methods?** - a set of general principles that describes how any investigation is done. The method follows a specific sequence as indicated below:

   Observation – describes and measures a data

   Hypothesis – possible interpretation of the data (must be stated in a way that it can be verified or tested (e.g. If I say that putting my wet clothes in a dryer, the clothes will be dry in 30 minutes. The statement can be tested or verified by anybody

Module 1 Unit 2: cont'd

Experiments- test that is done to verified the hypothesis

Theory – consistent data from experiments that confirms the hypothesis. Then the hypothesis becomes a theory

- **What is Technology?** When scientific information is converted into commercial use , then, the process is called technology

Module 1 Unit 1: Practice Questions

(1) Define the following terms:

      (1) Matter

      (2) states of matter

      (3) chemistry

(2) List 3 different branches of chemistry?

(4) Give 4 examples of how chemistry is related to other discipline of studies?

(5) List and explain the steps of a scientific method?

(6) Define the term technology?

(7) What is a chemical?

(8) What is a chemical reaction?

Module 2: Matter and Energy

In this module, we shall study the following topics:

- Classification of matter
- Difference between an element and a compound
- What a mixture is
- Types of mixtures
- States and physical properties of matter
- Physical and chemical changes
- Measurements such as temperature , units of temperature, conversion of temperature scales, normal body temperature, significant figures, accuracy, precision, boiling and freezing points
- Types of error: systematic and random
- Prefixes used Scientific notations

Module 2 Unit 1
Classification of matter, mixtures, physical and chemical properties and changes

- **Classification of Matter**: Matter is anything that has mass or occupy space (can, wood etc)
- It can be classified according to its compositions:
    - Pure substances – are matter with fixed compositions, viz
        - Elements: the simplest form of matter that has only one component (H_2O, O_2, Cu, Fe, Ag, H etc)
    - Compounds: 2 or more elements chemically combined together
- **What is the difference between an Element and a Compound?** – An element cannot be broken down into another substance whereas a compound can be chemically broken down to a simpler substance (NaCl → Na + Cl)
- **What is a Mixture?** – 2 or more substances/elements physically mixed together e.g. air ($N_{2+}O_2$), Coke, sand/rice, tea, ocean etc
- Characteristic composition of mixtures may vary e.g. brass = Cu + Zinc (composition of each element in brass may vary)
- **Method of Separation of a mixture:** Mixtures are separated by physical methods – filtration (filter paper, sieves, used for liquids), Distillation (2 miscible liquids e.g. water and alcohol, Boiling point difference is very important), Chromatography (uses differences in the rates of migration of components on a stationary and mobile phase)

Module 2 Unit 1 cont'd

Classification of matter, mixtures, physical and chemical properties and changes

- **Types of Mixture**: **Homogeneous** – composition is uniform throughout the bulk of the sample (air, sea water). If mixture is a liquid, it is called a solution

- **Heterogeneous**: composition is not uniform throughout the bulk of the sample (oil/water) probably due to partial solubility (chalk/water, colloids, suspensions)

Module 2 Unit 2

State and Physical Properties of Matter, Physical and Chemical Changes

- **States/Physical Properties of Matter**: One other way to be able to describe matter is to observe its physical properties.
- Physical properties can be described as attributes that can be observed or measured with out affecting the identity of the matter e.g. color of hair, shape of an animal, melting point(MP), boiling point (BP), weight, height etc
- When applied to living matter , it is called biodata or biometrics
- Matter is organized by its component elements. The organization of these elements gradually progresses from: loosely organized (air)→More organized (liquid)→Most organized (solid)
- Solid: has shape and volume with moderately organized molecule that is next to each other and held together by attractive forces. It forms a rigid structure which only allows vibrations of its components in a fixed position
- Liquid: has no shape of its own but takes the shape of the container holding the liquid. It has volume and mass
- Air: has no shape of its own but assumes the shape of the container holding it. It has volume and mass

Module 2; Unit 2 cont'd

- **Physical Changes:**
- Matter can undergo a physical change
- Its state/appearance will change but its components will remain the same
- **Example**: Water can exist as a solid (ice by freezing), Liquid (normal state), gas (steam or water vapor, by boiling). In all cases, its components (H and O) remain the same
- **Physical changes** are reversible . For example, evaporation, filtration, distillation are examples of physical processes.
- **Examples of Physical changes**:

  Change of State: Water to steam (boiling, evaporation)
  Change of Appearance: Salt dissolves in water
   Change of shape: Hammering Copper to different shapes
    Change of size: Crushing ice to smaller sizes

- **Chemical Changes**
- Ability to form a new substance with new physical and chemical properties
- Examples: Gari + hot water → Eba

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

$$2Fe + 3/2O_2 \longrightarrow Fe_2O_3 \text{ (Rust, brown color)}$$

$$H_2O/heat$$

Module 2 Unit 3
Measurements, Temperature, Types of Errors, Accuracy and Precision

Measurements

- People often are confused about the difference between *accuracy and* precision. *Accuracy* means how close a data is to its "true" value while *precision* means how close a repeated set of data are to each other.

- In another way we may say that a data is *accurate* if it has very small amount of *systematic* error (errors that can be determined, determinate error) while *precision* means a data has a small amount of *random* (indeterminate error) error. However, precision will lead to accuracy except if there is a deliberate bias in the way we do things (e.g. weighing an object in different ways for each measurement).

- Accuracy deals with how we calibrate the devices or methods of measurement while precision deals with the resolution and quality of the devices or methods of measurement. All measurements are approximations and we can only specify our *precision* (not accuracy) by one digit beyond the last reported digit (significant figures). For example, if we say the temperature of a piece of metal is 30°C, what we are saying is that the temperature lies within 29.5°C and 30.5°C

- **Temperature:** measures the degree of warmness of a substance

- **Units of Temperature**: Celsius (°C), Fahrenheit (°F), Kelvin/Absolute (K)

- **Example**: Body Temperature – Normal 36.5-37°C (96.8- 98.6°F)

- **Hyperthermia** (above normal, > 37°C) – Fever, increased temperature, pulse, respiration, perspiration. > 40°C/107°F may result into death

Module 2 Unit 3 cont'd

- **Hypothermia** (below normal, < 37ºC) – pale, cold, decreased pulse, respiration, sluggishness, unconsciousness, if T < 26ºC client may die
- Treatment: Increase oxygen and fluid intake
- Conversion Scale: Thermometer scale uses the properties of water to boil and freeze as a reference parameter for measuring temperature changes
- Freezing point(FP): 0ºC/ 32ºF   Boiling Point (BP) = 100ºC/212ºF
- For a Celsius and Fahrenheit thermometer: Uses glass column of the same length filled with mercury.
- The interval between the FP and BP on a Celsius thermometer = 100
- The interval between the FP and BP on a Fahrenheit thermometer = 180
- Since the length of glass column for both thermometers are the same, hence we
- can say that $\dfrac{ºC -0}{100} = \dfrac{ºF -32}{180}$
-
- $ºC = \dfrac{5(ºF-32)}{9}$
-
- ºF = 1.86ºC + 32 or (9/5)ºC + 32
- Kelvin Temperature: Coldest Temperature – -273ºC – absolute temperature which is Zero on Kelvin's Scale (0K)
- There are no negative values of temperature on the Kelvin scale and no degree sign on its units of temperature
- K= ºC + 273

Module 2 Unit 4

Significant Figures, Significant figures in Calculation, Reporting Actual Precision, Scientific Prefixes

- **Significant Figures**
- All measurements are approximations and we can only specify our *precision* (not accuracy) by one digit beyond the last reported digit (significant figures). For example, if we say the temperature of a piece of metal is 30ºC, what we are saying is that the temperature lies within 29.5ºC and 30.5ºC. In general, within any set of data, the more nonzero digits the more precision is implied. A scientist must set a realistic limit for its data. A realistic limit is set when the range of the data is not too large.
- Significant Figures (sf.): are non-zero (numbers) digits including those of decimal number
- 222.10 – 4 sf.
- 0.114560 – 5 sf.
- Zero may be a significant digit depending on its position within a number
- Zero in between non-zero digits is a significant digit
- 222.012 – 6 sf.
- 222.1014 – 7sf
- 0.110140 – 5 sf.
- Zero place immediately after a decimal point with no other number following is significant
- 25.0 – 3 sf.; 16.00 – 4 sf.
- Zero preceding a non zero digit in a decimal number is not significant: 0.001- 1 sf.

Module 2 Unit 4: cont'd

Significant Figures in Calculations

- **Multiplication:** The answer is rounded to the same number of significant figures as the number with the largest significant figure
- Example: Area = Length x breath. If Length = 2.2 cm and Breath = 1.12 cm
- Area = 2.2 x 1.12 = 2.464 = 2.46 cm^2
- **Division**: The answer is rounded to the same number of significant figure as the number with the least significant figure
- Example: $\dfrac{2.2 \text{ x } 1.12}{0.144}$ = $\dfrac{2.46}{0.144}$ = 17. 83333 = 17.3

Reporting Actual Precision

- If the digit next to the desired precision is ≥ 5, then add 1 to the digit before the number of desired precision (significant figure) and drop other digits
- Example: Round up 37.7863 to 4 sf. Answer: 37.79
- **Rounding off Numbers**: If the digit to be dropped is ≤ 5, then drop this number and others after this number
- Example: Round up 37.7823 to 4 sf. Answer: 37.78

Module 2 Unit 4: cont'd

- Prefixes: Can be placed in front of a unit to either increase or decrease its size by a factor of 10
- **Prefix**
- **Multiplication Factor (Increases by a factor of 10)**
- Peta (P) 10^{15}
- Tera (T) 10^{12}
- Giga (G) 10^{9}
- Mega (M) 10^{6}
- Kilo (K) 10^{3}
- **Prefix Multiplication Factor (Decreases by a factor of 10)**
- Deci (d) 10^{-1}
- Centi (c) 10^{-2}
- Milli (m) 10^{-3}
- Micro (μ) 10^{-6}
- Nano (n) 10^{-9}
- Pico (p) 10^{-12}
- Femto (f) 10^{-15}

Module 2 Practice Questions

(1) The three states of matter are ----------, ------------- and ------------

(2) Convert 212°F to K

(3) Rearrange the following process in descending order of occurrence: Experiment, Theory, Observation, Hypothesis

(4) A baby girl weighs 10kg at birth. The baby body temperature is 120°F two hours after birth. The nurse received an order from the physician to give the baby 5mL/Ib of paracetamol every eight hours for three days.

(a) Calculate the baby's body weight in pounds (Ib)

(b) How many mL of paracetamol will the baby receive every eight hours?

(c)   What is the baby's temperature in °C?

(d) Is the baby experiencing hyperthermia or hypothermia?

(e) Give one symptom the baby will show for this condition?

(5) An empty 25-mL cylinder weighs 60.75 g and weighs 86.05g when filled with water. When filled exactly with the same volume of ethanol, the weight is 80 g. Calculate the density of ethanol?

(6) Answer True or False

- Physical properties are those characteristics that can be observed or measured without affecting the identity of the substance

- A gas takes the shape of its container

- Matter can be converted from one form to another form

Module 3 Atoms and Elements

The following topics will be discussed in this module:
- Composition of matter, elements and atoms
- Atomic mass scale, isotopes
- Dalton's atomic theory, structure of an atom
- Atomic particles, sub atomic particles, arrangement of sub-atomic particles
- Atomic number and mass,
- Energy levels and arrangement of electrons in principal energy levels
- Number of electrons in each principal energy levels
- Electronic configuration
- Principles guiding the arrangement of electrons in sub atomic orbitals
- Electronic diagram and configuration

Module 3 Unit 1
Composition of matter, Elements and Atoms

- **Matter consist of** - Elements (simplest form of matter)
- There are 117 elements known for now e.g. H, Na, Cl, O_2, N, C etc.
- All known elements are arranged in increasing atomic number in the Periodic Table
- All elements are represented by a symbol (abbreviation that is used to represent an element e.g. Na, K
- Symbols consist of one or two letters – H, Na, N, Ca
- **Element:** are made up of atoms
- Atom is from the word ATOMOS – minute particles
- Atoms are the smallest particle or basic unit of an element that can take part in a chemical reaction
- It has the same characteristics as the element
- Each element has its own characteristic atom different from those of other elements
- Atoms are tiny particles not visible to the naked eyes but can be viewed by powerful microscope such as Scanning Tunneling Microscope (STM)
- For example, 1.0 g of Na and 16 g of Oxygen contain 2.6×10^{22} atoms and 6.02×10^{13} atoms respectively

Module 3 Unit 1 cont'd

- The mass of one atom of Na is 3.8175×10^{-23} and that of oxygen 2.66×10^{-23}
- These measurements were made with highly precise instruments such as a mass spectrometer and nuclear reaction energies

Atomic Mass Scale

- The mass of an atom is in the range of 1.0×10^{-23}. This is so small that it may not be useful or convenient in normal laboratory scale
- Physical scientist and chemist devised a relative scale which is more suitable for calculating atomic masses. This scale is called the **atomic mass scale (ams)**
- Atomic masses are measured relative to a reference element that is arbitrarily chosen for convenience sake and assigned an arbitrary number.
- The unit used for this scale is called the **atomic mass unit (amu)**
- Earlier Chemist chose naturally occurring isotopes of oxygen as the reference element and defined **amu** as 1/16 of the average mass atoms of isotopes of oxygen: ^{16}O, ^{17}O, ^{18}O
- Earlier Physicist chose the predominant isotope of oxygen: ^{16}O and defined **amu** as 1/16 of the mass of an atom of ^{16}O
- In 1961, and international consensus (SI unit) chose isotope of carbon-12 (**C-12**) as the reference element and arbitrarily assigned the value of 12 amu
- Therefore, **1 amu = 1/12 of mass of C-12 atom**

Module 3 Unit 1 cont'd

- Mass spectrometer gave the real mass of one atom of C-12 as **1.9927×10^{-23}g**

- Hence **1 amu = 1/12 x 1.9927×10^{-23}g = 1.6606×10^{-24}g**

- If mass spectrometer or other precise methods measures the mass of one atom of any element (e.g. one atom of Na weighs **3.8176×10^{-23}g),** one can then convert this to its equivalent on the **amu scale**

- On **amu scale**, the mass of one atom of Na will be **3.8176×10^{-23}g/1.6606×10^{-24}g = 22. 9892 amu**

- The Biologist uses the Dalton scale which is 1 amu = 1 Dalton

- **This is how the atomic masses of all elements are calculated and recorded in the periodic table**

- A simple way to explain the concept of **amu scale** is to consider a basket full of 20 cubes of sugar. Each cube of sugar has different weights. If we randomly choose one of the cubes (labeled X with mass 20g) as our reference cube to which the masses of all other cubes in the basket is calculated

- We also assign an arbitrary unit of 12 to cube X

- If we pick cube Y with mass 40g, we can calculate the relative mass of cube Y on our arbitrary scale as = **40/20 x 12 = 24 amu**

- We can calculate the relative masses of other cubes of sugar in the basket in the same way

- **In practice, the average mass of isotopes of elements is used in calculating is relative atomic or atomic masses**

- For practical purposes**, the gram-atomic mass unit** is used in laboratory calculations

Module 3 Unit 1 cont'd

- The **gram-atomic mass** is the amu expressed in grams and is defined as the molar mass of that element

- For example, the amu of Na= 23, the gram atomic mass (molar mass of Na) = 23g

- **Example:** The average relative mass of magnesium atoms is 4.03065×10^{-23}g. Calculate its atomic mass if one atom of C-12 has a mass of 1.9927×10^{-23}g and an atomic mass of 12.0 amu

- **Answer:**   First step is to compare the relative atomic mass of magnesium atoms with that of 1/12 of C-12

- 1amu of C-12 = $1/12 \times 1.9927 \times 10^{-23}$g = 1.6606×10^{-24}g

- Therefore, the atomic mass of the magnesium ions will be =

-      4.03065×10^{-23}g$/1.6606 \times 10^{-24}$g = 24.505 amu

Module 3 Unit 2
Isotopes, Daltons Atomic Theory, Structure of an Atom

- **Isotopes:** In nature, the same element might have 2 or more different atoms with different masses but with all other characteristics the same (e.g. atomic number). **Such atoms are called isotopes**.

- **Isotope is defined as atoms of an element with the same atomic number but different atomic masses or number.**

- Not all isotopes of an element are stable. Some are radioactive such as those of aluminum.

- Isotopes exist in different amount in nature. The amount of each isotope in nature is referred to as **its relative abundance**

- The relative abundance of any element is always constant. That is, if any sample of an isotopic element is analyzed, the amount of each isotopic atom will be the same.

- A good example is Chlorine: It has isotope ^{35}Cl and ^{37}Cl with atomic masses of 35 and 37 amu and relative abundances of 75.53 % and 24.47% respectively

- If an element M has three isotopes ^{a}A, ^{b}B, ^{c}C (a, b, c – respective atomic masses) and abundances of I%, J% and K% respectively,

- Atomic mass of M = (1/100) (Ia + Jb + Kc)

Example 1: For Chlorine, its atomic mass = (1/100) (75.53x35 + 24.47x37) = 35.5 amu

Module 3 Unit 2 cont'd

- **Example 2**- Element T has two isotopes T-35 and T-37. Calculate the natural composition of these isotopes, if the atomic mass of T is 35.5 amu

- Answer: Let x% represent the abundance of isotope T-35, then the abundance of isotope T-37 will be (100-x)%

- Therefore the atomic mass of element

- $T = 35.5 = 1/100 (35x + 37(100-x))$

- $x = 75$ for T-35 and $(100-75) = 25$ for (T-37)

Daltons Atomic Theory

- In the 1900's, Dalton improved on Democritus postulate that matter can be divided into small indivisible particles. He proposed the following

- All matter are made up of tiny particles called Atoms

- All atoms of the same element are identical and different from those of other elements e.g. atoms of Na are different from those of oxygen

- 2 or more atoms of different element combine together in a fixed ratio to form a new substance

- The combination is called a chemical reaction which consists of rearrangement, separation, combination, and formation processes.

- Atoms never created or destroyed (Law of conservation of matter: Matter is never created nor destroyed. It is converted from one form to another)
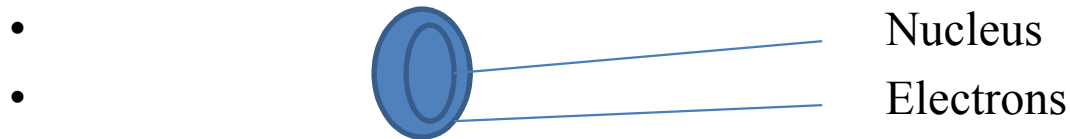
Module 3 Unit 2 cont'd

- **Structure of Atom**
- In 1897, J.J Thompson suggested that atoms can be divided to subatomic particles.
- He applied electricity to a glass tube (cathode ray tube) containing a positively charged electrode
- He observed that some particles where attracted to these positive electrode and called them cathode ray
- He concluded that these particles must be negatively charged because they are attracted to a positive electrode and therefore called them **Electrons**
- Since an atom is neutrally charged, he also suggested that there will be equal number of positively charged particles in the atom which he called **Protons**
- He finally suggested that the structure of atom is like a "pudding" with the electrons forming a cloud and the protons randomly scattered in the cloud
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- **Rutherford**: tested Thompsons atomic structure model
- He assumed that if a positively charged particle is passed through Thompsons atom, if should pass through it undeviated
- In his alpha particle scattering experiments, He passed positively charged particle through a thin gold foil

Module 3 Unit 2 cont'd

- He observed that some particles were slightly deviated, some moderately, some severely and then some totally reflected in the opposite direction.

- This observation contradicts Thompson model

- He also observed that this positive particles are heavier than the electrons and 10,000 times smaller in size than atomic diameter

- He concluded that for the deviations and reflections to occur, there must be a concentration of positive charges of the gold foil that repels these particles as pass near or through it. The extent of closeness of these positively charge particle to these concentrated centre will determine the extent of deviation or reflection

- The positive center he called **nucleus.**

- He then suggested a model of a positive nucleus surrounded by negatively charged particles (electrons)

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Nucleus

Electrons

Chadwick also concluded that atoms have some particles that have **mass but no charge**. These particles are called **neutrons**

Module 3 Unit 2 cont'd

- **Subatomic Particles:** In summary, an atom is said to contain a proton and neutron (called nucleons) that resides in the nucleus and surrounded by clouds of electron(s).

- The proton and the neutron have about the same mass

- Both are about 1840 times heavier than an electron.

- Details of the properties of each of these particles is described below

- Electron (e^-): is negatively charged (-1). Mass = 5.486×10^{-4} amu or 9.109×10^{-28} g, negative charge of -1.602×10^{-19} Coulombs.

- Proton (p): positively charged (+1). Mass = 1.0087 amu or, positive charge of $+1.602 \times 10^{-19}$ Coulombs

- Neutron (n): neutrally charged. Mass = 1.0087 amu or 1.6748×10^{-24} g

Module 3 Unit 2 cont'd

| Particle	Abbreviated Notation	Mass (amu)	Mass (g)	Relative Charge	Charge (coulombs)
Electron	e	Negligible (5.486×10^{-4})	9.109×10^{-28}	-1	-1.602×10^{-19}
Proton	p	1	1.6748×10^{-24}	+1	$+1.602 \times 10^{-19}$
Neutron	n	1		0	0

Module 3 Unit 3
Arrangement of Sub Atomic Particles, Atomic Number/Atomic Mass

Arrangement of Subatomic Particles
- **Rutherford's Planetary Model:**
- From experiments the following conclusions were reached about the structure of an atom
- Electrons resides outside the nucleus in shells of certain energy level
- Electrons removes randomly in an orbit around a positively charged nucleus and sustained in their orbit by both centripetal and centrifugal forces
- The nucleus contains only the protons and neutrons
- All atoms are neutrally charged and contains the same number or protons and electrons
- A neutron has no charge
- The mass of an atom is determined by the mass of the nucleus, that is, the mass of the proton and the neutron
- Isotopes of the same element occurs because of differences in the number of neutrons in the nucleus of its atoms
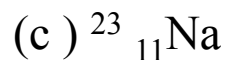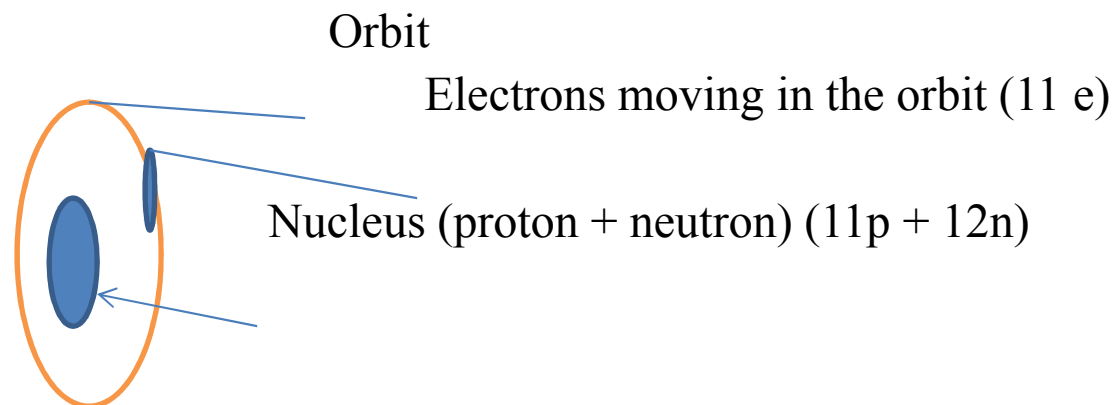
Module 3 Unit 3 cont'd

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- Orbit

- Electrons moving in the orbit

- Nucleus (proton + neutron)

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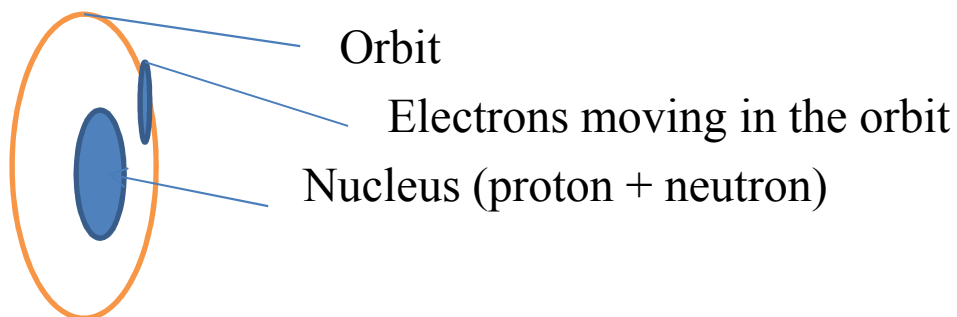
- Atomic Number/Atomic Mass

- All atoms of the same element have the same number of protons and electrons

- This number is called the **Atomic number (Z)**

- The periodic table is an arrangement of elements in increasing atomic number

- The mass of all atoms is the mass of its nucleus and is called the **Atomic Mass, A** and **$A = Z + n$**

- The **atomic mass** is written at the **top left corner** of the symbol that represents the element while the **atomic number** is written at the **left lower corner e.g.**

- $_{11}{}^{23}\textbf{Na}$

Module 3 Unit 3 cont'd

- Example: X is an element having 11 protons in its nucleus. Given that the relative atomic mass of this element is 23.
- Find
- Mass number
- Number of electron and
- Number of neutron the atom of the element
- Draw a simple illustration of the arrangement of the subatomic particles
- Write the symbol of the element indicating mass number and the atomic number
- Answer: (i) A = 23   (i) Z = 11 = the number of the protons
- (iii) A = Z + n therefore, 23= 11 +n   and   n = 23 − 11 = 12
-
-   Orbit

                                                Electrons moving in the orbit (11 e)

                                          Nucleus (proton + neutron) (11p + 12n)

-
- (c) $^{23}_{11}$Na

Module 3 Unit 4

Arrangements of Electrons in Principal Energy Levels

Arrangement of Electrons in Principal Energy Levels

- Neil Bohr – Postulated that electrons moves in a circular orbit or shell around the nucleus

- The orbits are at a fixed distance and energy level from the nucleus

- Electrons arrange themselves in this orbit or shell

- The energy level increases as orbit distances increases from the nucleus

- There is an energy value associated with each orbit or shell

- The maximum number of energy level for a known atom is seven and is called the **principal energy (quantum) level with a symbol n**

- The principal energy level is labeled 1-7 or alphabetically from K to Q (KLMNOQ)

- Each energy level have a maximum number of electrons that can occupy it and is defined as **$2n^2$** as shown in the table below

Module 3; Unit 4 cont'd

Number of Electrons in Principal Energy Levels

| Principal Energy Level	Alphabetical Label of Energy Level	Maximum Number of Allowed Electrons
1	K	2
2	L	8
3	M	18
4	N	32
5	O	50
6	P	72
7	Q	98

Module 3 Unit 5
Electronic Configuration

- **Electronic Configuration**

- Electronic configuration of an atom is written in accordance to the number of electrons found in each energy level available in the atomic structure of an element starting with the energy level numbered 1 or K

- For example $_{11}Na = 2, 8, 1$

- Example: Show diagrammatically how the subatomic particles are arranged in the following atoms. Also show the electronic structure for each of them.

- $_{11}{}^{23}Na$ (b) $_1{}^1H$ (c) $_8{}^{16}O$

- State the electronic configuration and the principal energy levels at which the outermost electron can be found in each atom, then calculate the maximum number of electron allowed in this principal energy level,

- The exact location of an electron cannot be identified. A region where there is a high probability of finding an electron is called an atomic orbital. This region of space is three dimensional and results into a specific shape for each type of orbital as shown below.

- s-orbital: the shape is spherical and the probability of finding an electron at the nucleus is high

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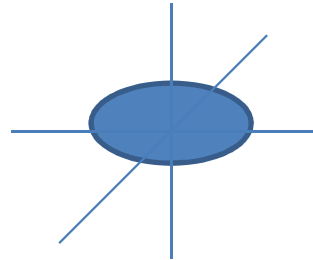
- p-orbital: consist of two adjacent spherical lobes positioned at opposite sides of a nucleus. The three degenerate p-orbitals lie along the 3-dimentional axes (p_x, p_y, p_z). The probability of finding an electron at the nucleus is zero. The nucleus of the p orbital is called a node point

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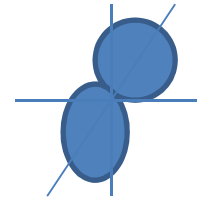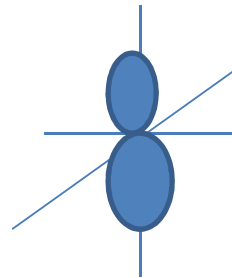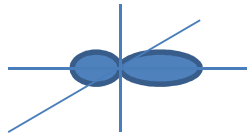
Module 3 Unit 5 cont'd

- The regions on both sides of the node points is assigned + and – signs. These signs do not represent ionic charges. The s orbital is assigned the + sign because it has no node point.
- The shapes of the d and f orbitals are complex and cannot be easily represented
- The five d degenerate orbitals are : d_x, d_y, d_z, d_{xy}, d_{xz}, or d_{yz}
- The seven f degenerate orbitals are: f_x, f_y, f_z, f_{xy}, f_{xz}, $f_{yz,}$ f_{xyz}

Module 3 Unit 6

Arrangement of Electrons in Sub-Orbitals, Electronic Configuration, Electronic Diagram

- **Electrons are arranged in sub-orbital by three principles, viz:**

- Aufbau or building principles: electrons are filled into orbitals in order of increasing energy- 1s<2s<2p<3s<sp<4s<3d<4p<5s<4d<5p<6s<4f<5d etc

- Pauli's exclusion principle: Electrons spins on its axis and hence generates magnetic and electric fields. Electrons in the same degenerate orbital repels each other and therefore will have opposite spins number to be able to reside in the same orbital. Hence Pauli formulate the following rules: (i)The maximum number of electrons in a degenerate orbital is two (ii) the two electrons in a degenerate orbital have opposite quantum spin number (+/-). Each spin is represented by half headed arrows (   ). (iii) no two electron have all four principal quantum numbers the same

- Hund's rule (the principle of maximum multiplicity): degenerate orbitals are singly filled before pairing up. As a result, an atom tends to have as many unpaired electrons as possible, In order words, electrons avoid repelling each other by seeking out empty shells to fill up instead of pairing up.

-

Module 3 Unit 6 cont'd

- The electronic arrangement can also be written in the format shown below by following the direction of the arrow

- 1s
- 2s   2p

- 3s   3p   3d

- 4s   4p   4d   4f

- 5s   5p   5d   5f

- 6s   6p   6d   6f
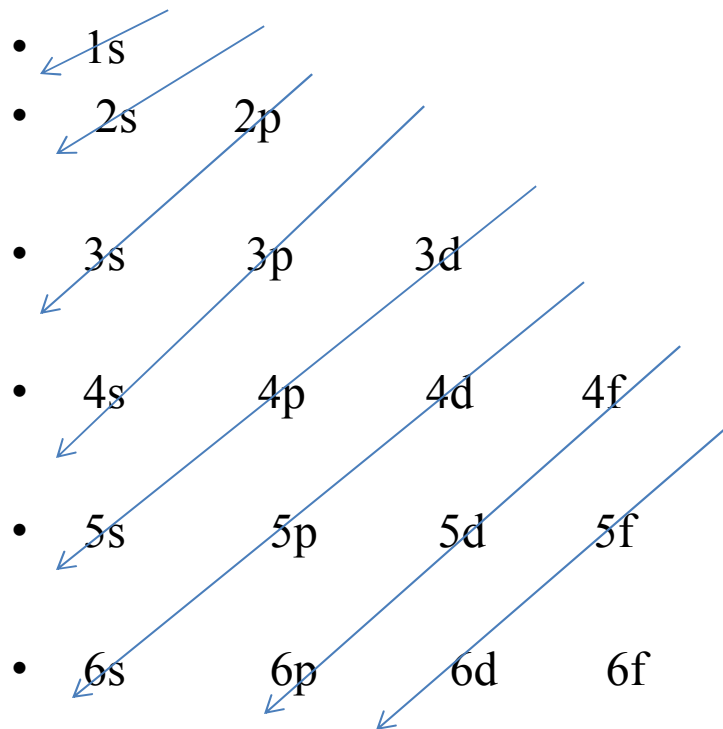
Electronic Configuration and Electronic Diagram

- The shorthand form of representing the arrangement of electrons in orbitals can be in form of a diagram or in a specified configuration.

- For example, carbon has 6 valence electrons and can be represented thus:

- Electronic Configuration: $1s^2\,2s^22p^2$

- The correct arrangement of the sub-orbitals is show in the previous slide

Electronic Diagram: The electronic diagram of carbon can be drawn as shown below

-

Abbreviated Format

- Another way of writing electronic configuration is to use the **Abbreviated format.** This involves using the configuration of noble gas preceding the desired element and filling the remaining valence electrons in the next principal quantum level

- For example, Sodium has 11 valence electrons: $1s^2\,2s^22p^63s^1$. The nearest noble gas before sodium is Neon with electronic configuration: $1s^2\,2s^22p^61s^2\,2s^22p^6$ which is one electron shorn of sodium configuration. Hence sodium configuration can be written as:$[Ne]3s^1$ which is the condensed format of sodium configuration.

Module 3 Practice Questions

(1) State Pauli's exclusion principle?

(2) Which of the following orbital arrangement is correct

  (a) 1s2s2p3s3d4s3p4p5s4d

  (b) 1s2s3p3s2p3d4s4p5s4d

  (c) 1s2s2p3s3p4s3d4p5s4d

(3) Write the abbreviated and the electronic configuration of Chlorine.

 (4) Answer True of False

  (a) Electrons are paired up one orbital at a time before the next orbital of the
      same energy is filled up

  (b) Electrons spin of a pair of electron in an orbital must have the same direction
       to form a stable element

  (c) Chlorine with mass 35 and 37 is an example of a neutron

  (d) Brass is made up of iron and copper

  (e) One cm^3 is the same as one milliliter

 (5) State two of Dalton Atomics theory?

 (6) Calculate the contribution of each chlorine isotope to the atomic mass of chlorine,
   if they occur in nature as 75.76% and 24.24% in nature respectively? (Chlorine
   isotopes are Chlorine 35 and 37)

Module 4: Periodicity

The topics included in this chapter include:
- The periodic table
- Periods and Groups
- Groups with special names
- Metals, Metalloids, Non-metals
- Physical Properties of Elements in the Periodic Table
- The Electronic Configuration and the Periodic Table

Module 4 Unit 1: Periodicity

- **The Periodic Table**
- History- 1800 – scientist noticed that some elements looks and behave alike
- 1872- Mendeleev and Meyer arranged 60 elements known then according to their properties and increasing atomic masses
- Today, we have 117 elements arranged according to increasing atomic properties or similarities in their electron distribution in their valence shells (s and p orbitals, orbitals that hold the electrons involved in bonding)
- For example, Oxygen, O([He] $2s^22p^4$) and Sulfur, , S([Ne] $3s^23p^4$) have same number of electrons in the valence shell and hence, have similar properties as well as some other differences (O is a colorless gas while S is a yellow solid both at room temperature)
- 1915- Moseley found that X-ray frequencies of elements increases with increasing atomic weight
- He assigned a whole number to X-ray frequencies of each element and called it atomic number which he assigned to the number of protons in a nucleus of an atom
- The atomic weight increases as the atomic number increases and he arranged the atomic number in increasing order as seen in the periodic table
- This arrangement lead to two trends in the periodic table: the period which is the horizontal row of the table and the Group; the vertical rows in the periodic table
- There are 7 periods (top to bottom) and 18 groups (left to right) in the Periodic Table (PT)

Module 4 Unit 1 cont'd

- **Period:**
- 1st period- 2 elements (H and He)
- 2nd Period – 8 elements, Li – Ne
- 3rd Period – 8 elements, Na – Ar
- 4th Period – 18 elements, K – Kr
- 5th Period – 18 elements, Rb- Xe
- 6th Period – 32 elements, Cs – Lu
- 7th Period – 31 elements, Fr- Lr
- Lanthanides (period 6) and Actinides (period 7) are placed at the bottom of the table so that everything can fit into a page
- Group: Groups 1, 2, 13 - 18 are called representative groups (Initially called groups 1A, 2A, 3A - 8A)
- Center – Transitional Group, 3-12 (old name 1B to 8B)

Module 4 Unit 1 cont'd

- **Classification of Groups**
- Some groups have special names
- Group. 1A – Alkali metals , Li, Na, K, Rb and Cs
- Physical properties – soft, shinning metals, good conductors of electricity and heat, low melting point, reacts violently with water, forms white substances with oxygen to form oxides
- Hydrogen is placed on top of this group but is not a metal
- Group. 2A – Alkaline Earth metals, Be, Mg, Ca, Sr, Ba, Ra. Shinny metals but less reactive compare to Grp 1A
- Group. 7A(17) – Halogens, F, Cl, Br, I, At. Very reactive (especially F and Cl) and forms compounds with most elements
- Group. 8A(18) – Noble gases, He, Ne, Ar, Kr, Xe and Rn. Unreactive

Module 4 Unit 2: Metals, Metalloids and Non-Metals

- Metals, Mettlloids and Non metals
- There is a zig - zag line that runs from group 13 to group 17.
- The line separates elements into Metals (on the left, except H) and non metals (on the right)
- 13 14 15 16 17 - Group
- B
- Si
- Ge As
- SbTe
- PoAt
- Metals- shinny solids at room temperature except mercury, can be stretched (ductile) and flattened (malleable), good conductors of heat and electricity, high MP than non metals. Metallic character increases from right to left and top to bottom of the PT
- Non Metals- not really shinny and some are liquids or gasses, poor conductors of heat and electricity, not malleable or ductile, low MP
- Metalloids- Elements located along the zigzag line- B, Si, Ge, Te, Po, At (except Al), can be metal or non metal and will display characteristics of both but not as good as either of the two, often called Semiconductors (can be modified to conduct or insulate)
- Transitional Elements- Elements in between metals and non metals (Grps. 3 -12)

Module 4 Unit 4: Physical Properties of Elements in the Periodic Table

- **Physical Properties of Elements in the Periodic Table**
- Effective Nuclear Charge: electrons are negatively charged and hence are attracted to the nucleus. Properties of atoms depend on how strongly they are attracted by the nucleus and also on the electron configuration. Coulombs law says that the strength of the force between two charges depends on the size of the charges on the charges and the distance between them. Hence the force of attraction on an electron by the nucleus depends on the net charge on the electron and the distance between the electron and the nucleus
- $F = Z/r$ (z is the total charge and r the distance). The large the r the less is F
- In another way we can vies each value of r as describing the shell in which the electrons reside. The force that keeps the electron in its orbit is
- $F = mv^2/r$ (mass of electron, v its velocity and r the radius of the orbit) (1).
- The Kinetic energy of each electron is described as, $KE = \frac{1}{2} mv^2$ (2)
- From (2) $v^2 = 2KE/m$ and
- substitute in (1) $F = m(2KE/m)$
- Therefore, $F = 2KE/r = K/r$ (where K= constant= 2KE)
- Hence, force of attraction on the electron varies with r i.e. it decreases with increasing r or vice versa. This variations determine how the physical properties of elements varies across the PT.
- In a multiple electron atom, the electrons experience both attractive force by the nucleus and the repulsive forces by the other surrounding electrons. On the average, each electron will experience the net electric field in the environment created by the nucleus and other electrons.

Module 4; Unit 3 cont'd

Property	Li	Na	K	Rb	Cs	Conclusion
Density g/cm^3 @ 20^{0}C	0.53	0.97	0.86	1.53	1.87	Increasing
Melting Point	179	97.8	63.7	38.9	28.3	Decreasing
Hardness	0.6	0.4	0.5	0.3	0.3	Decreasing
Electrical Conductance	17.4	35.2	23.1	13	8.1	Decreasing
Flame Color	Carmine	Yellow	Violet	Bluish red	Blue	

Module 4 Unit 3 cont'd

- We view the net electric field as if it resulted from a single proton in the nucleus and is called the **Effective nuclear charge (Z_{eff})** but in reality $Z_{eff} < Z$

Physical Properties

- **Ionization Energy**: Low ionization energy which decreases as we go down the group

- **Electron Affinity**: decreases down the group

- **Atomic size or radius**: Increases down the group

- **Metallic bonding**: Because of the large size and limited number of valence electrons, the metallic bonds between the atoms of these metals are not that strong. This also account for the low density, melting point and hardness.

Module 4 Unit 4
The Electronic Configuration and the Periodic Table

- **The Electronic Configuration and the Periodic Table**

- Different sections/block of the PT belongs to the s, p, d, f sublevels

- Hence, we can build the EC of an element by reading the PT in increasing atomic #

- **S- block-** includes Grps 1A and 2A (1 and 2) of the PT including H and He. This implies that their valence electrons is in the s- orbital. The period # to which the element belong to indicates which s-orbital where the electrons reside (2s or 3s etc)

- **P-block-** Include element in Grps 3A(13) – 8A(18). The period # indicates which p- level the valence electrons resides in (2p, 3p etc), each period has 6-p electons

- **D-block element-** Transitional elements (groups 3-12). Each period has 10-d electrons

- **D and F(Partially filled shells) –** Lanthanides

- **D and F shells-** Actinides

Module 4 Practice Question

(1) Which of the following elements has three valence electrons?

   A) lithium B) boron C) nitrogen D) more than one of the above

(2) The electron configuration for gallium is: $[_{31}Ga]$

   A) [Ar] $4s^24d^{10}4p^1$ B) [Ar] $4s^23d^{10}3p^1$ C) [Ar] $4s^23d^{10}4p^1$ D) none of these answers is correct

(3) Bohr's model of the atom was able to accurately explain:

- A) Why spectral lines appear when atoms are heated.

- B) The energies of the spectral lines for each element.

- C) Why electrons travel in circular orbits around the nucleus.

- C) None of the above answers is correct.

(4)What subatomic particle has a mass of one atomic mass unit?

- A) proton B) neutron C) electron D) more than one of the above

(5) How many electrons does iron have?$(_{26}Fe)$

- A) 26 B) 30 C) 56 D) It depends on the isotope of iron

Module 5: Electromagnetic Radiation

We shall discuss the following topics in this module:

- What electromagnetic radiation (EMR) is
- Different forms or types of EMR
- EMR as a form of energy
- Properties and components of EMR
- EMR described as a wave or a particle
- Characteristics of a wave and how they are related to each other
- Arrangement of different types of EMR according to increasing or decreasing wavelength or frequency (EMR spectrum)
- Common units used in EMR applications
- Particle nature of EMR
- Planks explanation of radiations emitted by solids: Quantum energy
- Photoelectricity and photons: Albert Einstein explanation
- Types of emissions of the EMR: line, band, continuous
- Bohr's planetary model of an atom and its limitations
- Energy levels in an atom: principal quantum energy levels
- Maximum possible number of electrons in a principal energy levels
- Wave Nature of EMR: De Broglie's principle and Heinsberg Uncertainty Principle

Module 5 Unit 1
Electromagnetic Radiation and Properties and Components of Electromagnetic Radiation (EMR)

- O Lord, How wonderful thou art……. When I look at the works of thy hands ……Ps 8
- These psalm magnifies the marvelous things that God has given to us for our use. Such things include the EMR
- The EMR includes radiations such as UV, IR, visible light, gamma, microwave etc.
- All are forms of electromagnetic radiations (EMR) with similar characteristics
- EMR possesses energy which it carries through space, it is called radiant energy
- All EMR unlike sound, does not need a medium for it to be transmitted. It can travel through a vacuum.
- All EMR moves through vacuum at the speed of light, 3×10^8 m/s (c) or 3×10^{10} cm/s. The speed varies from one medium to another
- EMR is a stream of discrete energy particles called photons
- The energies are proportional to the frequencies of its radiation, hence it can be described as a wave or as a particle
- EMR can be viewed as an electrical force associated with a magnetic force (at 90^0 to each other). Both forces are vector quantities (has magnitude and direction)
- Plotting of either of the forces as a function of time or distance gives a graphic representation of a beam of EMR

Module 5 Unit 1cont'd

- The electrical force of a radiation is more useful in analysis than the magnetic force because it is responsible for events such as: reflection, refraction, absorption and transmission of a radiation

- Viewing an EMR from one end makes it look like a circular cross section that fluctuate its radius periodically from zero to its maximum amplitude. A two dimensional view is shown in the figure below.

- The power P, of a radiation is the energy of the beam that reaches a given area per second.

- The intensity is the power per unit solid angle

- Power and intensity can be substituted for each other
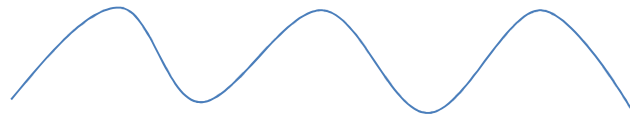
Module 5 Unit 2

• Wave Nature of EMR, Characteristics of a Wave, EMR Spectrum, Common Units used in EMR Applications

The Wave Nature of EMR

- All EMR can be described as a wave.

- The characteristics of a wave is as follow:

- Each wave is a packet of energy with high and low points
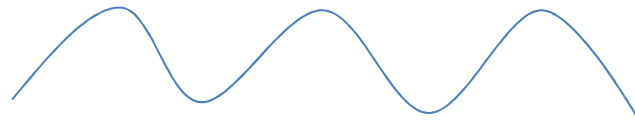
For example – when a stone is dropped in a pool of water or a boat is travelling on water, we observe some continuous ripples moving through the water. These ripples occur as a result of the energy transferred to the water by the stone or by the movement of the boat on the water. The energy moves through the water and is reflected as the ups and down movement that appears as ripples

Module 5: Unit 2 cont'd

- The highest point (maxima) of a wave is called the Peak (Crest) and the low point is a trough (inverted peak)
- The height of the peak is called the Amplitude
- The amount of energy possessed by a wave determines how high a peak is
- The distance between two successive peaks or trough is called the wavelength (λ called lambda). The unit of wavelength is mm, cm, m, km.
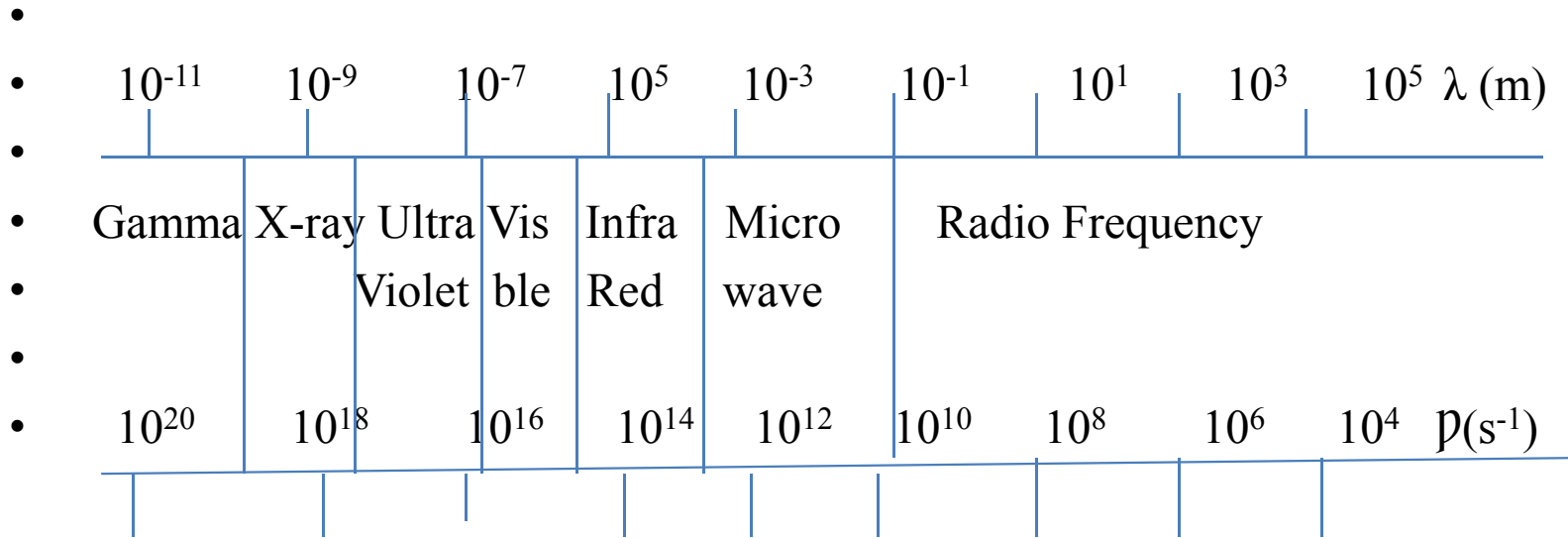-
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-
$$\lambda$$

amplitude
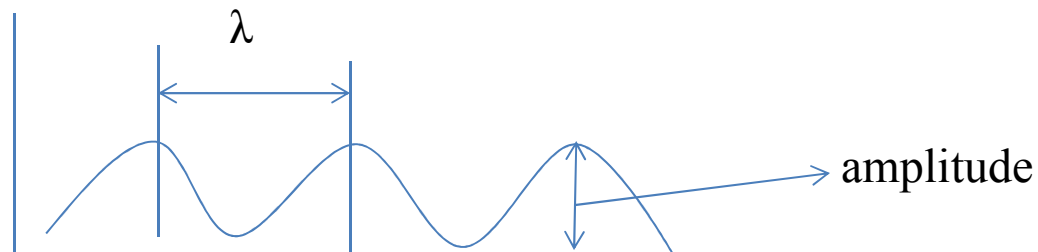
Module 5 Unit 2cont'd

- The time required for successive maxima to pass through a specific point is called the **Period**

- The number of wavelength or cycles that passes through a specific point per second is called the Frequency (ν called nu). The unit of frequency is Hz or s^{-1}

- All EMRs have electrical and magnetic fields

- The periodic oscillations of the intensities of both the electrical and magnetic fields results into the various characteristics of EMR waves

- The speed of water waves varies depending on the amount of energy input when it is created but that of EMR's are the same and equal to the speed of light

- Hence, there is a direct relationships between EMR parameters such as wavelength, energy, frequency etc For example, (i) $c = \lambda \nu$ (ii) if wavelength is long, the frequency is short (fewer cycles passing a point per second) and vice versa)

- The differences among EMR's is in their wavelengths or frequency

- We can arrange the different EMR's according to increasing wavelength or decreasing frequencies. This forms what is described as a continuous spectrum and is called the electromagnetic spectrum

Module 5 Unit 2cont'd

-

10^{-11}	10^{-9}	10^{-7}	10^5	10^{-3}	10^{-1}	10^1	10^3	10^5 λ (m)

-

- Gamma | X-ray | Ultra Violet | Vis ble | Infra Red | Micro wave | Radio Frequency

-

10^{20}	10^{18}	10^{16}	10^{14}	10^{12}	10^{10}	10^8	10^6	10^4 $\nu(s^{-1})$

- The EMR spectrum is very wide. Gamma rays is similar to the size of an atomic nuclei while that of Radio waves may be longer than a football field

- Human eyes can see in the visible light region of the EMR spectrum. However, this region is very narrow (400-700 nm)

-

Module 5: Unit 2 cont'd

- Table 3 Common Wavelength Units for EMR

Unit	Symbol	Length (m)	Type of Radiation
Angstrom	$\overset{o}{A}$	10^{-10}	X-ray
Nanometer	Nm	10^{-9}	UV-Visible
Micrometer	μm	10^{-6}	Infrared
Millimeter	Mm	10^{-3}	Microwave
Centimeter	Cm	10^{-2}	Microwave
Meter	M	1	TV, radio
Kilometer	Km	1000	Radio

Module 5 Unit 3

Particle nature of EMR, Planks Explanation of Radiations Emitted by Solids (Quantum energy), Photoelectricity and photons (Albert Einstein explanation)

Energy of Electromagnetic Radiation (Particle Nature of EMR): Quantized Energy and Photons

- Not all interactions of matter and EMR can be explained by the wave theory. It cannot explain the following phenomenon: (i) emission of light when an object is red hot (black body radiation) (ii) emission of electrons from the metal surfaces of which light shines (photo electricity) and (iii) emission of light from excited gas atoms (emission spectra)

- **Max Planck**: Solids emit radiation when heated. For example, metals glow red (kitchen stove) or emit white light (tungsten light bulb). A white hot object is hotter than a red hot object.

- Planck explained these phenomenon's by saying that energy is released or absorbed by atoms in discrete amount or in minimum size. This amount is called "QUANTUM" of fixed amount. He related energy to radiation as:

- $E = h\nu = hc/\lambda$

- (h is Planck's constant = 6.626×10^{-34} joule second, ν = frequency and E = energy of radiation)

- In summary, Planck is saying that energy is absorbed or released in multiple's amount of $h\nu$, $2h\nu$, $3h\nu$ etc.

Module 5 Unit 3 Cont'd

- If we go up a ramp, there is a gradual increase in potential energy (pd) while going up a stair case requires stepping on a stair one at a time which requires a fixed amount of energy (a fixed amount of pd rise per stair case, quantized) for each stair climbed. Each stair is regarded as an allowed action and therefore quantized energy levels are allowed energy states

- **The Photoelectric and Photons: Behavior of EMR as a Particle**
- Albert Einstein Theory:
- When light shines on a metal surface, the metal surface emits some electrons. This is called the photoelectric effect
- There is a minimum frequency of light required to cause this emission for different metals
- For example light with frequency > 4. 60 x 10^{14} s^{-1} will cause Cesium to emit light. Anytime less than this frequency will have no effect on Cesium.
- Einstein explained the observation as follows: Each stream of light behaves as a packet of energy or as a particle called photon
- Each photon has an energy $E = h\nu$ which is quantized and can be absorbed by electrons on the surface of the metal.

Module 5 Unit 3 cont'd

- Electrons are held together by attractive forces in the bulk of the metal and needed a minimum amount of energy called the work function to overcome these attractive forces.

- So, if the photon energy is greater than or equal to this work function, electrons will be emitted from the surface of the metals as observed with photoelectric effect. Excess photon energy will be converted to kinetic energy for the electrons

- A burst of light might be intense but if its frequency is less than the work function, it will not cause electron emission

Module 5: Unit 4

Types of emissions of the EMR, Bohr's planetary model of an atom and its Limitations, Energy levels in an Atom, Maximum number of electrons in a principal Energy levels

Emission of EMR

- Matter (atoms, ions, molecule) have fixed number of discrete and quantized energy levels
- Most specie of matter exist in the lowest energy level (ground state) at room temperature
- Matter can be excited to higher energy level by in many ways: heating (flame or arc), exposure to high voltage AC spark, intense source of EMR
- Life time of excited energy states are transient (10^{-6} – 10^{-9}s) and it will **relax** to lower energy level by emitting energy in form of radiation, heat or both
- Emission spectrum is often a plot of the power of radiation *vs*. wavelength or frequency
- There are 3 types of emission spectra: line, band, continuous which are useful in chemical analysis
- Line, Band, Continuous Spectra
- A source of radiant energy may possess a single wavelength like a laser light. This type of radiation is called a monochromatic ray of light.

Module 5: Unit 4 cont'd

- In nature, may radiant light like electric bulb, the sun, moon, and the stars emit light with many wavelengths. Also metals when heated to incandescence emit radiation of different wavelengths called black body radiation

- A **continuous** spectrum is produced when these wavelengths are separated into individual components that fade into each other. For example, if a light from a bulb is passed through a glass prism, the white light will be separated into its individual wavelength which appears as different colors that fades into each other

- Similar thing happens when the sun rays passes through raindrops in the atmosphere, the ray of light will be divided into different wavelengths which appears as different colors called the Rainbow

- **Band Spectrum**: Energy associated with each electronic energy state has associated quantized vibrational and rotational energy levels. The differences between these vibrational and rotational energy levels are much smaller than those between electronic energy levels and hence occur as bands of closely banded lines in between electronic energy levels. High resolution instruments can detect this band of spectra

Module 5 Unit 4 cont'd

- **Line Spectrum**: When electric current is passed through a tube containing different gasses at reduced pressure, each gas emits different colors characteristics of such gases.

- Similar thing happens to some metals. For example, Na emits yellow light; Neon emits red-orange color used as street lights etc.

- If these lights are passed through a prism, few distinct color lines are seen instead of a continuous spread of colors. These lines represent different wavelengths and are called line spectrum of those metals or gases

- For example, the hydrogen atom shows 4 lines: violet (410nm), blue(434nm), blue-green(486nm), and red (656nm). Other lines were also found in the UV and IR region of the hydrogen spectrum

- These lines fits into Balmer's series or the Rydeberg's equation

- $$1/\lambda = (R_H)(1/n^2_1 - 1/n^2_2)$$

- λ = wavelength, R_H= Rydberg's constant (1.096776×10^7 m^{-1}), n_1 and n_2 are integers with $n_2 > n_1$

- This explains why there are line spectra, that is ; energy is quantized

Module 5 Unit 4 cont'd

- **Bohr's Planetary Model**

- Rutherford suggested the planetary motion of electrons around the nucleus

- Classical physics says that such electrons will emit EMR and therefore will lose energy and therefore spiral down into the nucleus.

- Such phenomenon is not observed with atoms of any element

- Bohr suggested the following reasons for the electron not spiraling into the nucleus: (1) Only certain orbits of specific radii and with specified amount of energy are allowed for the electron of an hydrogen atom (2) An electron in an allowed energy state has a specific energy and is in a allowed energy state.

- An electron in an allowed state will not absorb or radiate energy

- An electron will absorb or radiate energy only if it moves from one allowed state to another allowed energy state.

- This energy is emitted or absorbed is a photon, $E = h\nu$

- Bohr's postulate is only true for Hydrogen atoms and not for other atoms because electrons behaves as a wave and it is supposed to spiral down into the nucleus (will lose energy)

Module 5 Unit 4 cont'd

- Bohr calculated this energy with the equation
- $E = (-hcR_H)(1/n^2) = -2.8 \times 10^{-18} J(1/n^2)$ [h = plank's constant, c = speed of light and R_H = Rydberg's constant, n is an integer = 1- infinity = principal quantum number]
- n = 1 is the lowest energy level (ground state of hydrogen) and as its value becomes larger, E, also becomes less negative and larger
- For n $\geq$ 2, the energy states are called Excited states
- When n = ∞ (infinity), E = 0 (zero energy) the electron will be removed from the orbit at this stage to another one (absorption or emission of energy)
- The value of the absorbed or emitted energy is equal to the difference in energy between the two states. $\Delta E = E_2 - E_1 = h\nu$
- Hence, it is only radiant light with frequency = ν that can be absorbed or emitted by an atom
- Since $\nu = c/\lambda$,
- then $\Delta E = h\nu = hc/\lambda = -2.8 \times 10^{-18} J(1/n^2_f - n^2_i)$ then
- $\lambda = hc/\Delta E$
- where n^2_f = final principal quantum state and n^2_i the initial principal quantum state
- If $n_f < n_i$, then ΔE = negative and hence energy is released (emission). The reverse will result into absorption

Module 5 Unit 4 cont'd

Limitations of Bohr's Model

- Bohr's model is only true for hydrogen atom
- It cannot explain why electrons do not spiral down into the nucleus if electron behaves like a wave. He just assumes it will not happen

Module 5 Unit 5

Wave Nature of EMR: De Broglie's principle and Heinsberg Uncertainty Principle

Wave Nature of Matter

- De Broglie said that an electron moving round a nucleus behaves like a wave because it has a certain wavelength associated with it
- The wavelength $\lambda = h/mv$

  where h = Planks constant, m = mass and v the velocity of the electron
- He proposed that any moving matter or particle have the same characteristic wavelength and a momentum, mv,
- De Broglie postulate was proved experimentally by passing a beam of electron through a crystal. The electrons were diffracted and magnified
- This led to the development of electron microscope which can magnify objects 3 million times more than a visible light can achieve
- The Uncertainty Principle
- If a ball is rolling down a ramp, we can accurately calculate its position, speed and direction at any time. But for an electron that behaves like a wave (wave extends into space) and a particle, its exact

Module 5 Unit 5 cont'd

Wave Nature of EMR: De Broglie's principle and Heinsberg Uncertainty Principle

- location and momentum at a point in time cannot be determined with certainty. This is what is called Heisenberg Uncertainty principle. It is only applicable to subatomic particles.
- He expressed the uncertainty as:
- $\Delta x . \Delta(mv) \geq h/4\prod$
- where x = position and mv= momentum of the particle
- $\Delta x \geq h/4\prod \Delta(mv)$.
- The only measurement with uncertainty is the speed of the electron (v) hence,
- $\Delta x \geq h/4\prod (m\Delta v)$.
- If we use 9.11×10^{-31} kg as the mass of a hydrogen electron and its speed as 5×10^6 m/s and uncertainty of 1% in its speed.
- Then , $\Delta x \geq (6.626 \times 10^{-34} \text{J-s})/4(22/7) (9.11 \times 10^{-31} \text{ kg })(0.01 \times 5 \times 10^6 \text{ m/s})$.
- $\Delta x \geq 1 \times 10^{-9}$m which is greater than the size of an atom
- Hence, we have no idea of where the electron is, which is the kernel of Heisenberg's Uncertainty principle

Module 5 Practice Questions

1. Example Two EMR waves are represented as shown in the figure below
   - (a)
   -
   -
   -
     (b)
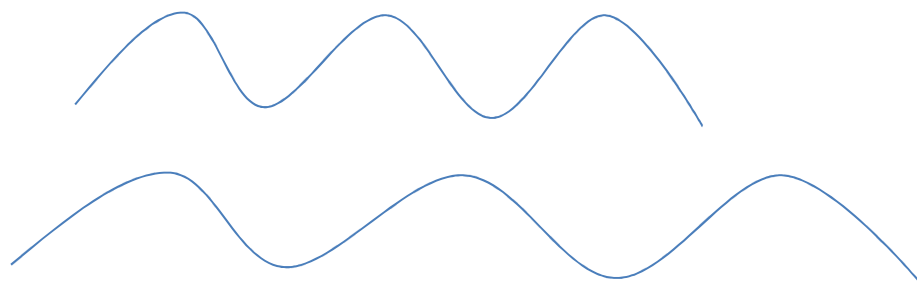
- (i) Which wave has the higher frequency (ii) If one wave represents the visible light and the other infrared radiation, which wave is which?

(2) The yellow light given off by sodium light used in public lighting has a wavelength of 589 nm. What is the frequency of this radiation?

(3) Assignment: (a) A laser light emits light with a frequency of 4.09×10^{14} s^{-1}. What is the energy of one photon of the radiation from the laser? (b) If the laser emits a pulse of energy containing 5.0×10^{17} photons of this radiation, what is the total energy of the pulse? (c) if the laser emits 1.3×10^{-2} J of energy during a pulse, how many photons are emitted during this pulse?

Module 5 Practice Questions

(4) If $E = (-hcR_H)(1/n^2_f - n^2_i)$, prove that $1/\lambda = (R_H)(1/n^2_f - 1/n^2_i)$ where

- λ = wavelength, R_H = Rydberg's constant n_1 and n_2 are integers numbers equivalent to the principal quantum numbers?

(5) (a) What is the wavelength of an electron moving at a speed of 5.97×10^6 m/s? (mass of electron = 9.11×10^{-31} kg).(b) In which region of the EMR can we find this wavelength?

Module 6 Inorganic and Organic Compounds

This module will discuss

- Definition and classification of a compound

- Ionic, covalent compounds and their structures

- The octet, quartet and duplet rules

- Formation of ions and ionic charges from group number

- Ions in the body and Food

- Formula Charge Balance in Ionic Compounds

- Naming and Writing Ionic Formulas

Module 6 Inorganic and Organic Compounds

- Naming Metals with Variable Charges
- Using Charge Balance to Write Formulas
- Polyatomic Ions (PAI)
- Formulas and Naming PAI's
- Covalent Compounds and their Names
- Electron Dot Formula of Covalent Bonds (Lewis Structure)
- Name and Formula of Covalent Compounds

Module 6 Unit 1
Definition and classification of a compound, Ionic, covalent compounds and their structures

- In nature, atoms are found in combination with other atoms. Only noble gases are unreactive
- **Compound**: are pure substances and are combinations of 2 or more element chemically combined together in a fixed ratio. They can either be ionic or covalent compounds
- **Ionic Compounds**: are formed when electrons are transferred from one metal to a non metal. Positive ions of the metals are formed while negative ions of the metal are also formed respectively. Both ions are bonded to each other by electrostatic attractive forces called Ionic bond
- Examples: $Na \rightarrow Na^+ + e$
- $\underline{Cl + e \rightarrow Cl^-}$
- $NaCl \rightarrow NaCl$
- Others are: baking powder ($NaHCO_3$), $CaCO_3$ etc
- Ionic compounds do not form molecules but form ions

- **Structures of Ionic Compounds**: Some ionic compounds can exist in more than one form e.g. Fe_2O_3 (ferrous oxide) or $FeSO_4$ (food supplement), SnF_2 (used in tooth paste to prevent tooth decay)

Module 6 Unit 1 cont'd

- Other examples include the different forms of Al_2O_3:
- (a) Sapphire (Blue color, has iron and titanium oxide as impurities) and (b) Ruby (Red in color with chromium as impurities.

- **Covalent Compounds**: are formed by at least two atoms sharing electrons between them to form a new bond.
- Covalent compounds form molecules.
- Examples of covalent compounds include compounds made of carbon and hydrogen which are called **organic compounds**
- Food molecules combine together to form bigger compounds called **biomolecules.** For example, Sucrose's (sugar) formula is $C_{12}H_{22}O_{11}$.

- **Module 6: Unit 2:The Octet, Quartet and Duplet rules, Formation of Ions and Ionic Charges from Group Number, Ions in the Body and Food**

- **Octet Rule**: Most elements combine together to form molecules or compounds.
- Noble gases rarely react because their outermost shell contains 8 electrons except Helium (He) which has 2 electrons. This configuration is called an Octet configuration and is a stable configuration
- Other elements react to form ionic or covalent compounds by loosing and gaining electrons or sharing electrons respectively with the aim of having 8 electrons in their valence shell (achieving an octet configuration)
- **Duplet State**: Only hydrogen gains an electron in covalent bonding to achieve a duplet state, i.e. have 2 electrons in its valence shell (also a stable configuration)
- **Quartet State**: Carbon has 4 electrons in its valence shell. This called a quartet state and it is a stable configuration
- **Formation of Ions**: Ions are formed when atoms loose or gain electrons
- Positive ions: Elements in groups I, II, and III easily looses electrons to achieve an octet configuration. In doing so, they form positive ions with a noble gas configuration (Octet configuration). These positive ions are called **Cations**

Module 6 Unit 2 cont'd

- For example, sodium EC $= 1s^2 2s^2 2p^6 3s^1$ looses the $3s^1$ electron to achieve the noble gas (Xe) configuration ($1s^2 2s^2 2p^6$)
- Na $\rightarrow$ Na$^+$ + e$^-$
- In general, metals loose their valence electrons to form positive ions.
- **Negative ions**: Non metals gain electrons to form negatively charge ions called Anions. In doing so they are able to achieve the octet configuration
- For example, chlorine has an EC $= 1s^2 2s^2 2p^6 3s^2 3p^5$ and will gain one more electron to achieve and EC $= 1s^2 2s^2 2p^6 3s^2 3p^6$
- Cl + e$^-$ $\rightarrow$ Cl$^-$
- **Names of Ions**: Names of anions always end with the suffix ' ide'. For example chlorine ion is called chloride ion. Others include bromide, fluoride, Iodide etc. Cations are called by their elemental names. For example, sodium ion will be called sodium ion.
- **Ionic Charges from Group Number**: Group numbers in the PT indicates the number of electrons in the valence shell of atoms
- It also represent the ionic charges of the ions of the elements in that group
- Metals in groups 1, 2 and 3 looses 1, 2and 3 electrons to form ions with +1, +2 and +3 ionic charges while non metals in groups 5A (13), 6A(16) and 7A (17) will gain 3, 2 and 1 electrons to form ions with -3, -2 and -1 ionic charges
- Group 4A(14) do not typically form ions because its valence shell contains 4 electrons which also is a stable configuration and does not easily loose electrons

Module 6 Unit 2 cont'd

Ions in the Body and Food

Ion	Occurrence	Function	Source	Hypo	Hyper
Na^+	Outside the cell	Regulate body fluid	Salt, tomatoes	Hyponatremia	Hypernatremia
K^+	Inside the cell	Regulate cell function	Bananas, nuts juice, orange	Hypokalemia	Hyperkalemia
Ca^+	Outside the cell	Tone/muscle function	Milk, dairy products	Hypocalcemia	Hypercalcemia
Mg^+	Outside the cell	Enzyme/muscle control	Spinach	Hypomagnesia	Hypermagnesia
Cl^-	Outside the cell	Gastric fluid regulation	Salt	Low chloride	High chloride

Module 6 Unit 3

Formula Charge Balance in Ionic Compounds, Naming and Writing Ionic Formulas, Naming Metals with Variable Charges

Ionic Compounds

- Ionic compounds can be positively or negatively charged
- The physical and chemical properties of these compounds are different from those of their corresponding elements
- For example
- Na +   Cl $\rightarrow$ NaCl
- Metal   liquid   white salt
- At room temperature- ionic compounds are stable
- ± ions attract and repel each other but result into 3D structure
- For example, NaCl structure includes 6 Na^+ ions is surrounded by 6 Cl^- ions and vice versa to form a cubic structure
- The force that holds these ions in place is called the electrostatic or ionic bond and it is a very strong bond
- Ionic bonding results into crystals with high melting point. For example NaCl melts at 501°C

Formula Charge Balance in Ionic Compounds

- Formula indicates the number of $\pm$ ions in an ionic compound

- The net charge on an ionic compound is always zero.

- Example is NaCl ➔ Na^+Cl^-

- Lewis diagram of an Ionic Compound

-                          ..   ..

-     Mg: + 2:Cl: ➔ $[Mg]^{2+}$ 2$[:Cl:]^-$

-                          .   ..

- **Naming and Writing Ionic Formulas**

- Ionic compounds with 2 elements- Name of metal is the same as the name of the element and the non metal is called by the first syllable of its elemental name ending with the suffix 'ide'. For example NaCl is called sodium chloride.

Module 6 Unit 4

Metals with Variable Charges, Naming Metals with Variable Charges, Using Charge Balance to Write Formulas, Polyatomic Ions (PAI), Formulas and Naming PAI's

- **Metals with Variable Charges-**
- Transitional metals forms 2 or more positive ions because they can lose their valence electrons and also electrons from their lower energy levels (p/d orbitals)
- For example Fe $\rightarrow$ Fe^{2+} + 2e$^-$ or Fe $\rightarrow$ Fe^{3+} + 3e$^-$
- We cannot predict ionic charges from their group number because of the possibility of forming variable charges
- Naming Metals with Variable Charges
- Write the elemental name first followed by the ionic charge in Roman Numeral in parenthesis e.g. Cu[II], Fe[III]
- Only Zn^{2+}, cadmium, Cd^{2+} do not have variable ionic charges among the transitional elements

Using Charge Balance to Write Formulas

- Example: write formula for Aluminum sulfide

-   Cation   Anion

-    Al   S

- Grp 3A(13)   Grp 6A(16)

- Al^{3+} S^{2-}

- To balance charges: Use the charge of the anion to multiply the cation and vice versa to balance the charges

-   $2Al^{3+}$ $3S^{-2}$

- Formula $= Al_2S_3$

Polyatomic Ions (PAI)

- Is a group of 2 or more atoms that has an ionic charge

- Most PAI consist of non metals such as P, S, C and N bonded to oxygen

- Most of the time they have ionic charges of -1, -2 or -3

- Example include NO_3^- (nitrate), SO_4^{-2}(sulfate)

- The only positive PAI is NH_4^+, the ammonium ion

Module 6 Unit 4 cont'd

- **Naming PAI**
- The name of the most common negatively charge PAI ends with the suffix 'ate' e.g. sulfate (SO_4^{-2})
- When PAI have one oxygen atom less than normal, its name ends in 'ite' e.g sulfite (SO_3^{-2})
- Others are: Nitrate (NO_3^-), Nitrite (NO_2^-), Carbonate (CO_3^-), Bicarbonate, HCO_3^-)
- Both forms have the same charges
- Elements in group 17 (7A) forms more than one type of PAI anions
- Prefixes are added to the names and the ending is changed to differentiates these anions
- Most common PAI ends with 'ate' e.g. Chlorate, ClO_3^-
- Example: PAI with the most oxygen more than the most common PAI have a prefix 'per'. Example include ClO_4^- is called perchlorate
- PAI with one oxygen less than the normal PAI ends with 'ite' e.g. ClO_2^- Chlorite
- PAI with 2 oxygen less than the most common PAI is called 'hypo' e.g. ClO^-, hypochlorite

Module 6 Unit 4 cont'd

- Formulas of PAI
- We use charge balance to write formulas for PAI. Example: $CaCO_3 \rightarrow Ca^{2+} + CO_3^-$; Aluminum carbonate is $Al^{3+} + 3HCO_3^- \rightarrow Al(HCO_3)_3$
- PAI are present in teeth and bones and is called hydroxyapatite, $Ca_{10}(PO_4)_6(OH)_2$

Module 6 Unit 5

Covalent Compounds and their Names, Electron Dot Formula of Covalent Bonds (Lewis configuration), Name and Formula of Covalent Compounds

Covalent Compounds and their Names

- When atoms of 2 or more non metals share electrons to form a bond between them, the bond is called a covalent bond. The atoms do not always share the electrons equally, so a polar covalent bond may be the result.

- Non metals have high ionization energy hence, they share electrons instead of transferring them as in ionic bonding

- Covalent bonding result into formation of a molecule.

- The simplest covalent molecule is the hydrogen molecule, H_2

- A hydrogen molecule is formed when 2 hydrogen atoms moves nearer each other. The nuclei of both atoms attract each other's electrons while their electrons repel each other. The force of attraction is greater than the force of repulsion and the net force makes both nuclei to share the electrons between them. Hence, both nuclei see the both electrons all the time. This arrangement makes both nuclei to achieve the noble gas configuration (octet or duplet configuration)

Module 6 Unit 5 cont'd

- Electron Dot Formula of Covalent Bonds (Lewis Structure)
- Lewis uses the chemical symbol of each element to represent its core electrons and then the dots to represent the valence electrons
- The hydrogen molecule will then be represented as
- **H· + H· → H:H**
-
- Sharing of electrons may occur between atoms of different elements to form a polar covalent molecule
-         ··   ··
- **H_x + :Cl: → Hx· Cl:**
-           ·   ··

Name and Formula of Covalent Compounds

- The chemical name of the metal is listed first
- Then chose the first syllable of the elemental name of the non metal and suffice it with "ide"
- Example: (i) CS_2 is named Carbon disulfide (ii) CCl_4 – Carbon tetrachloride

Module 6: Practice Questions

- 1 (a) Use one sentence to describe how an ionic compound is formed?

-

- (b) Give two examples of an ionic compound?
- (c) The force binding atoms (ions) together in an ionic compound is called----?

-

- 2. Both Sapphire and Ruby are Aluminum oxide. What is the difference between
- the two compounds?

-

- 3. (a)Two hydrogen atoms reacted with each other to form a hydrogen molecule. What type of bond exists between the two atoms of hydrogen? (b)What is the difference between this type of bond and an ionic bond?

Answer True of False

- Covalency involves donation of electrons
- Potassium is the only ion that is found inside animal cells
- Number of moles of a concentrated solution and the diluted one are different
- ClO_2^- is called chlorite ion and ClO^- the hypochlorite ion

Module 7 Unit 1
Types of gases, Physical Properties of Ideal Gases, Pressure

Different Types of Gases

- Gas is a form of matter that we are all familiar with. The most common one is Air which is a mixture of different gases such as: Nitrogen (N_2 78%, Oxygen (O_2, 21%) and other minor gases such as argon (Ar, 0,9%), water vapor etc
- Different gaseous substances have different chemical properties but all have similar physical properties
- Some gases exist as elements under normal temperature and pressure (NTP): H_2, He, Ar, Kr, Xe, N_2, F_2, Cl_2, O_2)
- Some gases exist as molecules: methane (CH_4), Hydrogen Chloride (HCl), Ammonia (NH_3) etc
- All gases are non metals with simple molecular formula and low masses
- Some solids (by sublimation, Iodine, camphor) or liquids (by vaporization, water) may also transform to gases and are often referred to as vapors

- **Physical Properties of Ideal Gases**
- Gases consist of small particles (atoms, molecules)
- The atoms or molecules of gases are very far apart from each other and they move randomly with high velocity

Module 7 Unit 1 cont'd

- Volume of gases are small compared to the volume of its container and hence it can expand and be compressed
- Gases may (ammonia, hydrogen sulfide) or may not have odor (oxygen) or color
- Pressure, temperature and volume changes affect gases
- Gases can mix together to form a homogeneous mixtures
- Mixtures of gases still retains their individual physical or chemical properties

- **Pressure**
- Pressure (P) is an amount of force (F) that moves a substance over a certain distance or acts on a certain area (A) in a certain direction.

$$P = F/A$$

- **Atmospheric Pressure**: Every form of matter experiences a force that tries to pull or push them towards the earth (center of the earth).
- For example, when a piece of mango is ripe, it falls to the earth, when we trip on an object we fall down (towards the earth). The pull towards the earth is called the gravitational force (9.8×10^3 atm or 9.8 m/s^2)
- **Gravitational force**, forces air to press down on the earth. This pressure is called the **atmospheric pressure**. At higher altitude the atmospheric pressure decreases

- The SI unit of force is the Newton (N) = $1 kgm/s^2$
- The SI unit of pressure is the **_pascal_** (Pa).
- Pascal = 1 Newton of force per meter squared $(1\ N/m^2) = 1\ kg\ m^{-1}s^{-2}$

- **Other Units of pressure include:**
- bar   1 bar 10^5 Pa
- atmosphere   1 atm   101.325 kPa or 760 mmHg
- torr   1 Torr $(101.325/760) = 133.32$ Pa
- milli meter of mercury   1 mmHg   133.32 Pa
- pounds per square inch   1 psi   0.894757 kPa

- **Example 1.** A substance has a mass of 1.0 kg and it exerts a pin point force on an area with size $1.0\ x\ 10^{-2} mm^2$ on the surface of the earth. (hint: force = mass x force of gravity, F= mg, g = $9.8\ m/s^2$). Calculate the pressure exerted the substance

- **Answer: P = F/A = mg/A** = 1.0 x 9.8 x $10^{-9}/0.01 = 9.8$ x $^{10-7}$ kg/ms^2

Example 2: Convert 0.480 atm. to torr?

$1 atm = 101325$ Pa

0.480 atm. $= 101325 \times 0.480$ Pa $= 48636$ Pa

$133,32$ Pa $= 1$ torr

Therefore,

48636 Pa $= 48636/133.32$ torr $= 364.81$ torr

The Gas laws

The Gas Laws

- There are 4 physical conditions that affect the properties of gases: Pressure **P**, Volume **V**, Temperature **T**, and amount of gas (number of moles, **n**)

- The equation that relates these parameters to each other is call the gas law

- Pressure-Volume Relationship: Boyle's Law

- If we inflate a balloon (expands), the volume inside it increases but the pressure expressed on it by the atmosphere decreases. The converse is also true

- **This is the Boyle's law**

- The Boyle's says: the volume of a fixed quantity of gas at constant temperature in inversely proportionally to it pressure

$$P \, \alpha \, 1/V \text{ or } PV = \text{constant}$$

- The value of the **gas constant** depends on the **amount of gas and its temperature**

-

Module 7 Unit 2 cont'd

V

1/P

P

1/V
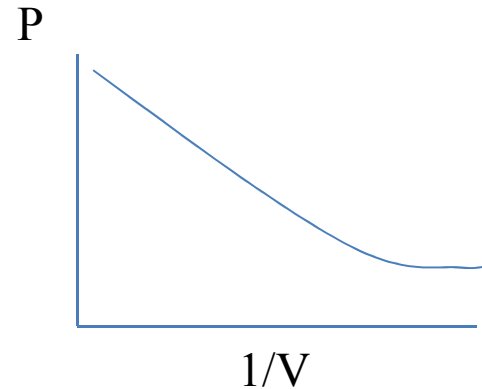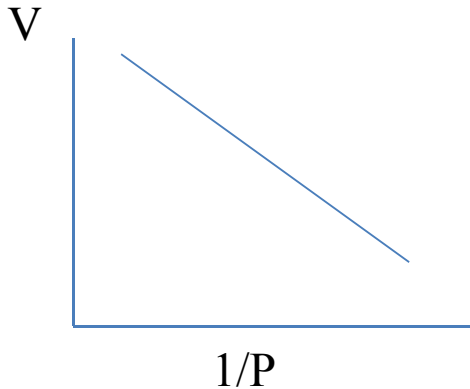
- Boyles law is demonstrated in our breathing mechanism: the diaphragm controls the volume of the lungs,

- Inhalation: the lungs expand (breathing in, V increases) while the rib cage also expands and the diaphragm stretches downward (pressure inside the lungs decreases.

- Exhalation: the lung is compressed (V decreases) while the rib cage is compressed and the diaphragm contracts upward (pressure inside lungs increases). The body temperature and amount of gas are constant for these two processes

Module 7 Unit 2 cont'd

Temperature-Volume Relationship: Charles Law

- Air expands as it is heated at atmospheric pressure. As temperature of air is increased, it becomes less dense, hence hot-air balloon rises at it is heated at atmospheric pressure. As the temperature of air is decreased, it becomes denser, hence if hot-air balloon is cooled it start to descend

- Charles law: shows a direct relationship between T and V and it says: the volume of a fixed amount of gas at constant pressure is directly related to its temperature.

- T α V that is V/T = constant   if amount of gas and its pressure are constant

- T α P that is T/P = constant   if amount of gas and Volume are constant

-

-

- V

-

-

- -300   300

- T⁰C

-
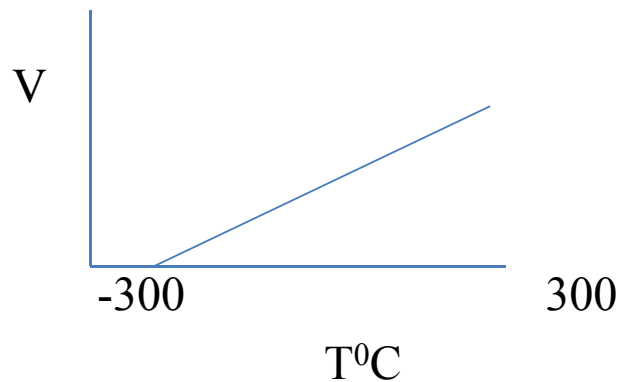
Module 7 Unit 2 cont'd

- The extrapolated V-T curve intersect the T axis at -273.15^0C where the volume of the gas is Zero. This temperature is never achieved because gases turns to liquid or gases before reaching this temperature

- This intersection is called the absolute temperature or absolute zero on a Kelvin scale

The Quantity-Volume Relationship: Gay Lussac and Avogadro's Law

- Gay Lussac law of Combining Volume: At a given pressure and temperature, the volumes of gases that react together are in a ratios of small whole numbers

- $2H_{2(g)} + O_2 \rightarrow 2H_2O$

- 2 volume of hydrogen gas reacts with one volume of oxygen gas to form 2 volumes of water

- Avogadro's Hypothesis: Equal volume of gases at the same temperature and pressure contain equal number of molecules

- 22,4 L of any gas at 0^0C and 1 atmosphere contain 6.02×10^{23} gas molecules which is 1 mole of that gas

Module 7 Unit 2 cont'd

- Avogadro's Law: The volume of gas maintained at constant temperature and pressure is directly proportional to the number of moles of the gas
- V α n i.e. V/n = Constant
- If volume is halved the number of moles of the gas will be halved and vice versa
-

Module 7 Unit 3
The Ideal Gas Equation and Standard Temperature and Pressure

The Ideal gas Equation

- We can combine together the Boyles, Charles and Avogadro's law to form what is called the ideal gas equation.

- V α 1/P (constant T, n) - Boyle

- V α T   (constant, P, n) - Charles

- V α n   (constant P, T) – Avogadro

- Combining the three laws gives: V α nT/P

- which translate to V = nRT/P where R is a constant called ideal gas constant

- PV= nRT - the general or ideal gas law (equation)

- The units of R depend on the units of P (atms, Pa, kPa), V(L), T (K), and n (moles)

- Other units can be used for pressure and Volume

- In using the ideal gas equation, the units of R must tally with the units of other parameters

Module 7 Unit 3 cont'd

The Units and Numerical Value of the Gas Constant R

Units	Numerical Value of the Gas Constant, R
L-atm/mol-K	0.08206
J/mol-K	8.314
Cal/mol-K	1.987
m^3-Pa/mol-K	8.314
L-torr/mol-K	62.36

Standard Temperature and Pressure (STP)

- If we have 1 mole of a gas at 0^0C and 1.0 atm. The volume of an ideal gas is given as
- V= nRT/P = (1.0 mole)(0.8206 L-atm./mol-K)(273.15 K)/1.0 atm.= 22.4 L
- 22.4 L is the volume of a mole of gas at STP

Module 7 Unit 4

Gas Densities and Molar Mass and Gas Mixtures and Partial Pressures

Gas Densities and Molar Mass

- We can use the ideal gas equation to calculate the density and molar masses of compounds
- $PV = nRT$ can be rearranged to give
- $n/V = P/RT$. If we multiply both sides with the molar mass M
- $nM/V = PM/RT$
- nM = mass of the substance in grams and density is mass/V
- hence, nM/V = density D
- Therefore, $nM/V = D = PM/RT$
- Implication of the equation: If M and P increases, density increases while increase in T indicates decrease in D
- We can also calculate the Molar Mass from this equation if other parameters are known

Gas Mixtures and Partial Pressures

- Daltons Law of Partial Pressures: The total pressures of a mixture of gases equals the sum of the pressures that each gas will would exert if it were present alone.
- $P_T = P_1 + P_2 + P_3\ldots\ldots$
- If all gases in this mixture obeys the ideal gas equation, then
- $P_T = (n_1 + n_2 + n_3\ldots)RT/V$

Module 7 Unit 5
Deviation from the Ideal Gas Behavior (Van der Waals forces)

Deviations from the Ideal Gas Behavior: Van der Waals Equation

- Real gases do not behave exactly like an ideal gas

- If we rearrange the ideal gas equation as n = PV/RT and make n = 1 for all gases (PV/RT = 1), then a plot of PV/RT versus pressure should be constant (equal to 1) at all pressures

- But there is a deviation at high pressures for all gases when this PV/RT is plotted against Pressures. However, at low pressures, the gases approaches the ideal behavior

- There is also a deviation at different temperatures. However, at high temperatures the plot approaches the ideal situation. At low temperatures, the deviation becomes very significant especially at the point of conversion to a liquid

- Because real gases have real volumes and their molecules interact with each other, significant deviations from the ideal behavior is usually very pronounced when defining P, V, n, R, T relationships

Module 7 Unit 5 cont'd

- Van der Waals made corrections to the ideal gas equation to accommodate all these interfering parameters by introducing two constants, a and *b*, where *a* accommodates the attractive forces between the molecules of the gases and *b* accommodates the finite volumes of the gases

- The ideal gas equation then becomes:

- $$(P + n^2a/V^2)(V + nb) = nRT$$

Module 7 Practice Questions

(1) If the volume of a fixed amount of gas is halved what will happen to the its temperature if pressure and its number of moles are kept constant

(2) If 1.00 mole of an ideal gas were confined to 22.4 L at 0.0^0C, it would exert a pressure of 1.00 atm. Use the van der Waals equation to estimate the pressure exerted by 1.00 mole of Cl_2 gas in 22.4 L at 0.0^0C (R= 0.08s06 L-atm./mole, a = 6.49 L^2-atm/mol^2, b = 0.0562 Lmol.)

(3) Define (i) vaporization (ii) sublimation (iii) Pressure

(4) State 5 properties of an ideal gas?

(5) State 4 physical parameters that affects the properties of gases?

(6) State (i) Boyles (ii) Charles (iii) Gay Lussac (iv) Avogadro (v) Dalton's law of partial pressures?

(7) Using the Boyles , Charles and Avogadro's laws; prove that PV= nRT

(8) What are the factors responsible for the deviation from the ideal gas behavior?

Module 8 Thermochemistry

This module discusses the following topics of Thermochemistry:

- Energy
- Heat and Units of Energy
- Specific Heat
- How to calculate specific heat
- Measuring Heat Changes
- Energy and Nutrition
- Energy Values of Food

Module 8 Unit 1
Thermochemistry: Energy, Heat and Units of Energy

Energy

- Energy is what we use when we perform work or other activities.
- It can exist in different forms such as heat, electricity, kinetic (to cause motion), potential (due to position or chemical composition of a substance) etc.
- Example of a KE is a car in motion and that of PE includes a boulder rolling down a hill or a city water tank on top of a platform
- We use energy when we walk, talk, climb, run, think, exercise emotions, eat, think etc
- When we are tired, it indicates that we do not have enough energy in our system to do any work
- Sources of energy: varies but includes the food we eat, fossils, fuel etc
- Energy can be converted from one form to another (Law of conservation of energy)

Heat and Units of Energy
- The SI unit for heat energy is the Joules (J) or Kilojoules (KJ)

Module 8 Unit 1 cont'd

- **Joules (J)** is a small amount of energy, so, the KJ is most often used unit for Heat energy. For example we need 10^5J to sleep for one hour which amounts to 100 KJ or we need to use 75000 J to heat a cup of tea (75KJ)

- The **Calorie, C**, is another unit of heat and is defined as the amount of heat needed to raise 1g of water by 1^0C.

- 1C= 4.184J

- **Example**: 1.0 g of propane fuel was used to cook dinner. 48000 J (48 KJ) of heat was released in that process. Calculate this amount of heat energy in calories

- **Answer**: 1 C = 4.184 J, then 48000 J = 48000/4.184 cal = 11, 000 C

Module 8 Unit 2
Specific Heat and How to calculate specific heat

Specific Heat

- Every substance has the ability to loose or gain heat energy whenever there is a change in temperature

- The amount of heat absorbed or released to reach a certain temperature varies from one substance to another one

- This capacity to absorb or release energy is a physical property and it is called the **Specific Heat (SH)**

- **SH** is defined as the amount of heat(q) needed to raise the temperature of 1.0 g of a substance by 1^0C

-     SH = Heat (q)/mass x ΔT = q/m x ΔT (J(Cal/g^0C)

- **The SH of water = 4.184J/ g^0C or 1 cal/ g^0C**

- 4.184 J of heat will raise 1g of water by 1^0C

- The same amount of heat will raise 1 g of copper by 10^0C and 1g of Al by 5^0C

- This indicates that water has a higher capacity than Cu or Al and also it has a higher capacity to absorb or release heat, therefore, it keeps the body temperature constant

- Low amount of SH for Cu and Al relative to water shows that they transfer heat more efficiently than water hence they are used as cooking wares and are better heat transfer substances

Module 8 Unit 2 cont'd
Specific Heat and How to Calculate Specific Heat

Specific Heat

- Every substance has the ability to loose or gain heat energy whenever there is a change in temperature

- The amount of heat absorbed or released to reach a certain temperature varies from one substance to another one

- This capacity to absorb or release energy is a physical property and it is called the **Specific Heat (SH)**

- **SH** is defined as the amount of heat(q) needed to raise the temperature of 1.0 g of a substance by 1^0C

- $SH = Heat\ (q)/mass\ x\ \Delta T = q/m\ x\ \Delta T\ (J(Cal/g^0C)$

- **The SH of water = 4.184J/ g^0C or 1 cal/ g^0C**

- 4.184 J of heat will raise 1g of water by 1^0C

Module 8 Unit 2 cont'd

Calculating Specific Heat (SH)

- What is the SH of lead if 530 J is added to raise the temperature of 35.6 g of lead by 12.5 ^{0}C?

- Answer: SH = q/m ΔT

-                 = 530/35.6 x 12.5 = 0.128 J/g^{0}C

- We can calculate amount of heat (q) or the mass of a substance by the following equation

-        Q = SH x m ΔT J and m = q/SH x ΔT

Module 8 Unit 3

Measuring Heat Changes

Measuring Heat Changes

- A calorimeter is used to measure temperature changes when a substance looses or gain heat
- A laboratory scale bomb calorimeter can be made with a Styrofoam cup (reactor) placed in another (bigger) Styrofoam cup (insulator)
- One may lag the insulator cup with glass wool
- Place a certain amount of water in the reactor and measure its temperature (Initial temp, T_1).
- Heat a certain object to a known temperature and drop it inside the water and immediately cover the calorimeter.
- Stir the water continuously and measure its temperature at specific intervals
- Stir until its temperature is constant and record it as the final temperature T_2
- Measure heat change by using the formula $\Delta T = T_2 - T_1$
- Heat lost by the object is equal to the heat gained by the water ($q_{object} = q_{water}$)
- $Q = SH \times m \times \Delta T$

Module 8 Unit 4
Energy and Nutrition and Energy Values of Food

Energy and Nutrition

- The food we eat produces energy for us to do work. Our body grow, sleep, respires etc as a result of energy supply from the food we eat.
- There are different types of food and all of them possesses a measure of energy
- Types of food include: Carbohydrate (main source of energy), fats and proteins (secondary source of energy if carbohydrate is not available)
- Food energy is measured in calories or kilocalories
- In nutrition unit, the calorie is written with a capital Cal. This implies that 1000 Cal = 1 Kcal
- We can also use KJ in nutrition unit. Remember that 1 Cal= 4.184KJ= 4184J
- For example, baked potato = 120 Cal = 4.184 x 120KJ ~ 500 J

Module 8 Unit 4 cont'd

Energy Values of Food

- Energy values of food are Kcal or KJ obtained by completely burning of 1g of food

Food Item	KJ/g	Kcal/g
Carbohydrate	17	4
Fats	38	9
Proteins	17	14

Module 8 Practice Questions

(1) 1.0 g of iron has temperature of 21^0C. What is the final temperature of the iron sample when 4.184 J is added? (SH_{iron} = 0.492 J/g^{0}C, SH_{water}= 4.184 J/g^{0}C)

(2) Define (i) the specific heat capacity (ii) calorie

(3) 1.0 g of propane fuel was used to cook dinner. 48000 J (48 KJ) of heat was released in that process. Calculate this amount of heat energy in calories

(4) Which instrument can be used to measure specific heat capacity of a substance

(5) If you don't have a bomb calorimeter in your laboratory, what can you do to measure the specific heat capacity of a metal object?

(6) Describe how you will use the equipment described in question 71 to measure specific heat capacity

(7) How can you determine the specific heat capacity of 1.0 g of yam?

Module 9
Thermodynamics

The following topics are included in this module:
- The first law of thermodynamics
- The second law of thermodynamics
- The third law of thermodynamics
- Zeroth Law of Thermodynamics

Module 9 Unit 1
Laws of Thermodynamics

- There are four **laws of thermodynamics** which define the fundamental physical quantities such as temperature, energy and entropy that characterizes thermodynamic systems.
- The laws describe how these quantities behave under various circumstances, and prohibit certain phenomena such as perpetual motion to occur

First Law of thermodynamics

- The first law of thermodynamics provides the basic definition of thermodynamic energy (which is also called internal energy) associated with all thermodynamic systems. It explains the rule of conservation of energy in nature.
- The law states that energy can be transformed, i.e. changed from one form to another, but cannot be created or destroyed.
- Because energy is conserved, the internal energy of a system changes as heat flows in or out of it. Equivalently, machines that violate the first law (perpetual motion machines) are impossible. Heat is the flow of thermal energy from one object to another.

Module 9 Unit 1 cont'd

- For example, wood (stored chemical energy) can be burnt (heat) and used to boil water. The heat energy from the boiling water can be use to drive a turbine to produce electricity (electrical energy). The electricity can be used to lighten a bulb (luminous energy) or play a radio (sound energy) etc

- The example shows that energy cannot be created or destroyed, it can only be converted from one form of energy to another form of energy

- The first law can also be formulated by stating that the change in the internal energy of a system is equal to the amount of heat supplied to the system, minus the amount of work performed by the system on its surroundings

$$\Delta U_{system} = Q\text{-}W$$

- For a thermodynamic cycle of a closed system, which returns to its original state, the heat Q_{in} supplied to a closed system in one stage of the cycle, minus that Q_{out} removed from it in another stage of the cycle, equals the net work done by the system.

$$\Delta U_{system \ (full \ cycle)} = 0 \text{ because } Q = Q_{in} - Q_{out} = W$$

- The increase in internal energy of a closed adiabatic system can only be the result of the net work performed by the system, because $Q = 0$.

$$\Delta U_{system \ (full \ circle)} = U_{final} - U_{initial} = \text{-}W$$

Module 9 Unit 2
The Second Law of Thermodynamics

Second Law of Thermodynamics

- The concept of energy in the first law does not account for the observation that natural processes have a preferred direction of progress. For example, spontaneously, heat always flows to regions of lower temperature, never to regions of higher temperature without external work being performed on the system.

- The first law is deals completely with moving from an initial state to a final state of an evolving system.

- The second law refers to a wide variety of processes, reversible and irreversible. All natural processes are irreversible. Reversible processes are a convenient theoretical fiction and do not occur in nature. A prime example of irreversibility is in the transfer of heat by conduction or radiation

- The second law of thermodynamics asserts the irreversibility of natural processes, and the tendency of natural processes to lead towards spatial homogeneity of matter and energy, and especially of temperature. It can be formulated in a variety of interesting and important ways.

- The explanation of this phenomenon can be made through the second law of thermodynamics

Module 9 Unit 2 cont'd

- A new physical property called the **entropy** is used by the second law of thermodynamics to explain this phenomenon
- A change in the entropy (S) of a system is the infinitesimal transfer of heat (Q) to a closed system driving a reversible process, divided by the equilibrium temperature (T) of the system.

$$dS = dQ/T$$

- For a closed thermodynamic system, it is a quantitative measure of the amount of thermal energy not available to do work.
- The second law also talks about the kinds of irreversibility other than heat transfer, such as friction, viscosity, and those of chemical reactions.
- The notion of entropy is needed to provide a wider scope of the law
- **One can also look at entropy from different angles such as:**
- A measure of the disorder or randomness in a closed system.
- A measure of the loss of information in a transmitted message.
- The tendency for all matter and energy in the universe to evolve toward a state of inert uniformity.
- Inevitable and steady deterioration of a system or society.

Third Law of Thermodynamics

- **The third law of thermodynamics**: The entropy of any pure substance in thermodynamic equilibrium approaches zero as the temperature approaches zero. The entropy of a system at absolute zero is typically zero, and in all cases is determined only by the number of different ground states it has.

- It is sometimes stated as follows:

- *The entropy of a perfect crystal of any pure substance approaches zero as the temperature approaches absolute zero.* At zero temperature the system must be in a state with the minimum thermal energy. This statement holds true if the perfect crystal has only one state with minimum energy. Entropy is related to the number of possible microstates according to:

$$S = k_B \, \ln\Omega$$

  where S is the entropy of the system, k_B Boltzmann's constant, and Ω the number of microstates (e.g. possible configurations of atoms). At absolute zero there is only 1 microstate possible ($\Omega=1$ as all the atoms are identical for a pure substance and as a result all orders are identical as there is only one combination) and $\ln(1) = 0$.

- It may also be stated as: The entropy of a system approaches a constant value as the temperature approaches zero. The constant value (not necessarily zero) is called the residual entropy of the system.

Module 9 Unit 4
Zeroth Law of Thermodynamics

- Zeroth Law of thermodynamics: If two systems are in thermal equilibrium with a third system, they must be in thermal equilibrium with each other.
- This law helps define the notion of temperature

Module 9 Practice Questions

(1) What is the first law of thermodynamics?

(2) Give an example that can demonstrate the first law of thermodynamics?

(3) Explain the second law of thermodynamics with examples?

(4) Describe entropy in four different ways

Module 10 Chemical Equilibrium

Topics that will be discussed in this module include:

- Rates of Reaction
- Activation Energy (AE)
- Measurement of Rates of Reaction
- Factors Affecting Rate of Reaction
- Equilibrium Constant
- Using K_{eq} to Predict Direction of Reaction
- Le Chatellier's Principle
- Effect of Change in Concentration
- Effect of a Catalyst
- Changes in Volume and Pressure
- Changes in Temperature
- Equilibrium Constant in a Saturated Solution
- Adding a Common Ion

Module 10 Unit 1
Chemical Equilibrium

Introduction

- There are different types of reactions: displacement, substitution, complex reaction etc.

- Some reactions are slow (aging of metals), some are very slow (reactions of medications in human body) and some are fast (acids reactions with carbonates, dissolution of sodium in water)

- The question then is: Why are reactions fast or slow?

- How fast or slow a reaction will go depends on many factors such as concentration, temperature, pressure etc

Rates of Reaction

- Rates of reaction (ROR) is a measure of how fast a reaction will occur

- Reactions happen when molecules, elements or atoms collide with the proper orientation and sufficient energy

- But not all collisions result into a reaction

Module 10 Unit 1 cont'd
Chemical Equilibrium

Activation Energy (AE)

- When the orientation of reactants are okay, there is still a need for enough energy to break the bonds between the atoms of each reactant for a reaction to occur

- Activation energy is the minimum amount of energy needed to break the bonds between atoms of reactants.

- AE is like going from one side of a hill to the other side. We need enough energy to get to the top of a hill and then can easily run down the hill to the other side

-   Activation Energy

-

-

-

Energy

- A + B   AB

- Reactant   Product

- Progress of reaction
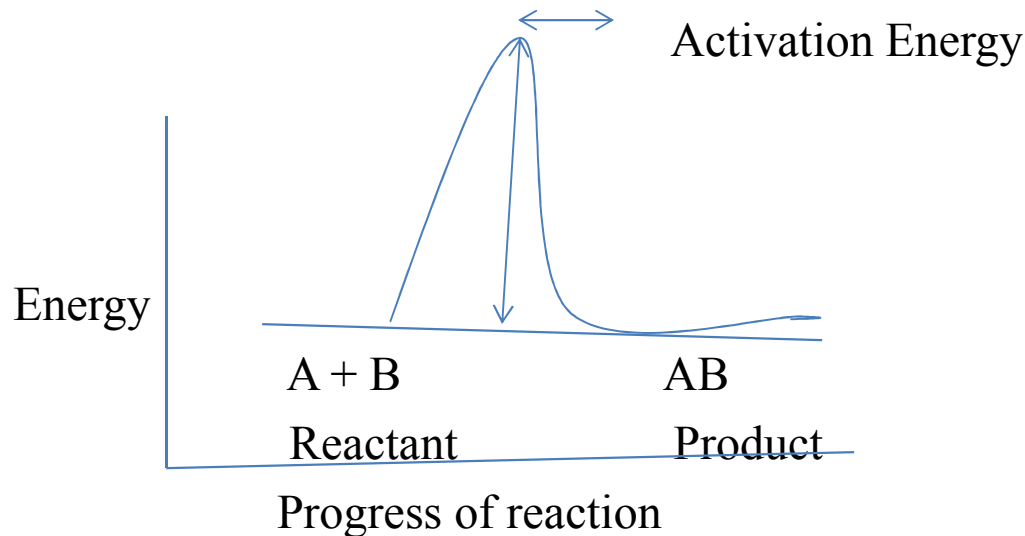
- Hence condition for a reaction to occur includes: collision, proper orientation, concentration and sufficient energy

Measurement of Rates of Reaction

- We can we measure the rates of a reaction by measuring the amount of reactants consumed or amount of product formed over a given amount of time

- R = [Δ in concentration]/Δ in time

- For example, we can buy a loaf of bread cut into 20 slices. We can measure how long it will take us to eat the whole 20 slices of bread. The we can calculate how long it took us to eat each slice of bread, which indicates the rate of reaction

- Example:

- (a) Distinguish between reactions rate and reaction mechanism (b) The following data were obtained for the reaction: I^-_{aq} + OCl^-_{aq} → IO^- + Cl^-_{aq}

	$[I^-]$	$[OCl^-]$	Initial Rate
	mol/dm^3	mol/dm^3	$mol/dm^2/s$
I	1.8×10^{-3}	2.1×10^{-3}	5.4×10^{-5}
II	3.6×10^{-3}	2.1×10^{-3}	10.8×10^{-5}
III	1.8×10^{-3}	6.3×10^{-2}	4.9×10^{-4}

- (i)What is the order of the reaction with respect to $[I^-]$ and $[OCl^-]$

- (ii) Calculate the rate constant using the results of experiment II

- (iii) What is the rate law or rate equation for this reaction?

Answer

- (a) reaction rate is the velocity with which reactants are used up (disappearance) to form products or the velocity at which products are formed while reaction mechanism is the pathway by which reactants are converted into products

- (bi) Rate = $k[I^-][OCl^-]$ Using experiments I and II

- $5.4 \times 10^{-5} = k[1.8 \times 10^{-3}]^x[2.1 \times 10^{-3}]^y$ (1)

- $10.8 \times 10^{-5} = k[3.1 \times 10^{-3}]^x[2.1 \times 10^{-3}]^y$ (2) Divide equation (2) by (1)

- $10.8 \times 10^{-5}/(5.4 \times 10^{-5}) = k/k (3.1 \times 10^{-3}/1.8 \times 10^{-3})^x(2.1 \times 10^{-3}/2.1 \times 10^{-3})^y$

- $2 = 2^x$ hence x = 1 and therefore, the order wrt. I^- is 1

- Using equation I and III

- $5.4 \times 10^{-5} = k[1.8 \times 10^{-3}]^x[2.1 \times 10^{-3}]^y$ (1)

- $4.9 \times 10^{-4} = k[1.8 \times 10^{-3}]^x[6.3 \times 10^{-2}]^y$ (2). Divide equation 2 by 1

- $4.9 \times 10^{-4}/5.4 \times 10^{-5} 5.4 \times 10^{-5} = k(1.8 \times 10^{-3}/1.8 \times 10^3)^x(6.3 \times 10^{-2}/2.1 \times 10^{-3})^y$

- $3^2 = 3^y$ and y = 2, therefore, the order wrt. $[OCl^-]$ is 2

Module 10 Unit 2 cont'd

Factors Affecting Rate of Reaction

- **Temperature**- increase in temperature increases the rate of a reaction by increasing the rate of collision, kinetic energy and energy. For example, increasing the heat when we are cooking increases the rate of cooking also if body temperature increases, the pulse rate and breathing will also increase. If we decrease temperature, ROR will also decrease. For example, freezing food decreases the rate of spoilage

- **Concentration of Reactants**- Increase in the concentration of reactants increases the rate of collision, hence the rate of reaction. Increase in **pressure** and **volume** also increases the rate of a reaction. The converse is also true.

- **Catalyst**: increases or slows down the rate of a reaction by lowering AE or providing and alternative pathway with a lower AE. So, if AE is lowered, more collision will occur, hence sufficient energy will be provided to make the reaction occur. For example, Platinum and Palladium are catalytic converters that convert pollutants form combustion of petrol in an automobile engine into harmless products. Pollutants in combustion engine include: CO, C_8H_{18} and NO

- $\qquad$ Pt

- $2CO_g + O_{2g} \rightarrow 2CO_{2g}$

- $\qquad$ Pt

- $C_8H_{18g} + 25\ O_{2g} \rightarrow 16CO_{2g} + 18H_2O_l$

- $\qquad$ Pt

- $2NO_g \rightarrow N_{2g} + O_{2g}$

Equilibrium Constant

- Reaction of reactants goes in forward and backward reaction at the same time and rate. At a certain time, the rate of the forward reaction and backward reaction will be the same. Then, the reaction is said to have reached equilibrium. Hence the [Reactants] and [Product] are constant

- If $aA + bB \leftrightarrow cC + dD$

- $$K_{eq} = [C]^c[D]^d/[A]^a[B]^b$$

- For example, $H_{2g} + I_{2g} \leftrightarrow 2HI_g$

- $K_{eq} = [HI]^2 /[H_2][I_2]$

- Some reactions are heterogeneous, i.e. the reactants and products are in more than one states in a reaction. The K_{eq} for the reaction is called heterogeneous K_{eq}

- The concentration of solids and liquids reactant or product are not included in the equation for equilibrium constant. For example, in the reaction shown below,

- $CaCO_{3s} \rightarrow CaO_s + CO_{2g}$ the $K_{eq} = [CO_{2g}]$

- In the reaction of $CO_{2g} + 2H_{2g} \rightarrow CH_3OH_l$ the

- $K_{eq} = 1/[CO_2][H_2]^2$

- If we know the concentrations of the reactants and the products, we can calculate the K_{eq} for any reaction

Module 10 Unit 3 cont'd

- For example, in the reaction of $H_{2g} + I_{2g} \leftrightarrow 2HI_g$ with the following concentrations:
  0.1M   0.2M   1.0M

- $\qquad K_{eq} = [HI]^2 / [H_2][I_2]$

- $\qquad\quad = [1.0]^2/[0.1][0.2] = 0.4$

Using K_{eq} to Predict Direction of Reaction

- If $A_g \leftrightarrow B_g$, $K_{eq} = [B]/[A]$

- If K_{eq} is large, then $[B] >>>> [A]$, the more product will be formed by the time equilibrium is reached. Hence the direction of reaction is towards the right. The converse is also true

- Le Chatellier's Principle

- Changes in conditions of reactions such as P, C, T, when a reaction is already in equilibrium, changes the position of equilibrium

- Le Chatellier's says that when a stress (change) is placed on a system already at equilibrium, the equilibrium will change (shift) in the direction that relieves the stress

- If a system is in equilibrium and a factor of equilibrium such as concentration is changed while others remain constant, K_{eq} will not change but the concentration of the other components and the direction of equilibrium will change so as to keep K_{eq} constant

Module 10 Unit 3 cont'd

- **Effect of Change in Concentration:** For example: $PCl_{5g} \leftrightarrow PCl_{3g} + Cl_{2g}$ $K_{eq} = [PCl_3][Cl_2]/[PCl_5] = 0.42$

- If $[PCl_5]$ is changed by adding mor PCl_5, while P, T are constant, the concentrations of PCl_3 and Cl_2 will change, but K_{eq} will not change but moves in the direction of the product.

- If $[PCl_3]$ and $[Cl_2]$ are increased, equilibrium will not change but shift towards the PCl_5 while its value remains constant

- **Effect of a Catalyst**: Addition of a catalyst to a reaction either fastens or slows down the rate of a reaction

- **Changes in Volume and Pressure:** Volume and pressure are inversely related to each other. But changes in their value will only change the direction of equilibrium and not its value

- **Changes in Temperature:** When Temperature is increased, the KE increase in

- the forward direction

- If the reaction is endothermic and the temperature is increased, the K_{eq} will increase and more products will be formed, hence the direction of equilibrium will be towards the right. If the temperature is lowered, equilibrium will shift to the left and the reaction will become exothermic

Module 10 Unit 4

Equilibrium Constant in a Saturated Solution

- In a saturated solution, some undissolved reactants are in contact with the
- maximum number of dissolved solutes. At equilibrium, the rate of dissolution will be equal to the rate of crystallization at a fixed temperature
- $CaC_2O_{4s} \leftrightarrow Ca^{2+}_{aq} + C_2O_4^{2-}_{aq}$
- Solubility constant of CaC_2O_{4s}, $K_{sp} = [Ca^{2+}_{aq}][C_2O_4^{2-}_{aq}]$
- Solubility of a solid is constant and is not included in the K_{sp}
- For example: $MgCO_{3s} \leftrightarrow Mg^{2+}_{aq} + CO_3^{2-}_{aq}$ $K_{sp} = [Mg^{2+}][CO_3^{2-}]$
- The mole ration of $Mg^{2+}/CO_3^{2-} = 1:1$ hence their concentration is the same, hence $K_{sp} = [s][s] = S^2$ and $S = \sqrt{K_{sp}}$
- If $K_{sp} = 1.6 \times 10^{-8}$, then $S = \sqrt{1.6 \times 10^{-8}} = 1.3 \times 10^{-4}M$

Adding a Common Ion

- For a sparingly soluble salt like $MgCl_{2s} \leftrightarrow Mg^{2+}_{aq} + 2Cl^-_{aq}$
- Adding Mg^{2+} which is soluble will increase the amount of Mg^{2+} ion solution, K_{sp} will shift direction to keep its value constant by making more Mg^{2+} reacting with more Cl^- to form $MgCl_2$ hence equilibrium will shift to the left
- This implies that the solubility of a sparingly soluble salt will decrease if a common ion is added.

Module 10 Practice Questions

- Define (i) Rates of a reaction (ii) activation energy

- How does equilibrium constant responds to the change in concentration of Chlorine in the reaction shown below: $PCl_{5g} \leftrightarrow PCl_{3g} + Cl_{2g}$

- If the equilibrium constant of a reaction is written as shown below, write the equation for the reaction: $K_{eq} = [HI]^2 /[H_2][I_2]$

- Write the equilibrium equation for the reactions shown below

- $CaCO_{3s} \rightarrow CaO_s + CO_{2g}$

- $CO_{2g} + 2H_{2g} \rightarrow CH_3OH_l$

- State Le Chatellier's principle?

- A loaf of bread consists of 20 slices. It took Nwachukwu 15 minutes to eat the 20 slices. What is the rate at which Nwachukwu ate the bread

- State 4 parameters that affects the rate of a reaction. Briefly describe how each one of them affects the rate of a reaction

- If a reaction is expressed as shown below, write an equation for its equilibrium constant

- If $aA + bB \leftrightarrow cC + dD$

Module 11 Acids and Bases

We shall discuss the following topics in this module:

- What acids and Bases are
- Different definitions of an acid and a base
- What are the properties on an acid and a base
- What conjugate acids and bases represent
- Strength of acids and bases

Module 11 Unit 1
Acids and Bases

Acid

- **Arrhenius Acid:** A substance that dissolves in water to give H^+ ions

- $HCl_l \leftrightarrow H^+_{aq} + Cl^-_{aq}$

- **Naming of Acids:** Acids dissolve in water to give H^+_{aq} ions and a negatively charged ion which is a non metal (Cl^-_{aq})

- The word **hydro** is placed before the name of the non metal and ends with the suffix 'ic' acid. For example, HCl is called hydrochloric acid, H_2SO_4 is called sulpuric acid

- Acids turns litmus paper red and phenolphthalein colorless

- Acids are sour to taste

- **BrØnstead Acids**: is defined as a substance that donate protons. H^+ ion does not exist freely in water. It attaches itself to a molecule of water to form the hydroxonium ion: $H^+ + H_2O \rightarrow H_3O^+$

- **Lewis Acid**: is a substance that accepts a lone pair of electrons in forming a coordinate covalent bond

Module 11 Unit 1 cont'd

Bases

- **Arrhenius Base:** a substance that dissolve in water to give a cation and a hydroxide ion. For example, NaO → Na^+_{aq} + OH^-_{aq}. Most Arrhenius bases are groups 1A and 2A substances e.g Na, K, Be, Li

- Bases like $Ca(OH)_2$, $Al(OH)_3$, $Fe(OH)_3$ are strong but insoluble in water

- Arrhenius bases are sour to taste and slimy to touch. They turn litmus blue and phenolphthalein pink

- **Naming of Bases:** The name of the metal is ended with the word hydroxide. For example NaOH is called sodium hydroxide.

- **BrØnstead Base:** is a substance that accepts a proton. For example, water accepts a proton to form the hydroxonium ion, hence water is a base in this reaction. Another example is the reaction: NH_3 + H^+ ↔ NH_4^+, NH_3 is proton acceptor and hence it is acting as a base in this reaction

- **Lewis Base:** is a substance that donates a lone pair of electron to form a coordinate covalent bond

Module 11 Unit 2

Conjugate Acid and Base

BrØnstead- Lowry: Conjugate pairs such as A-B are molecules or ions that gain or lose H^+ ions:

-
-
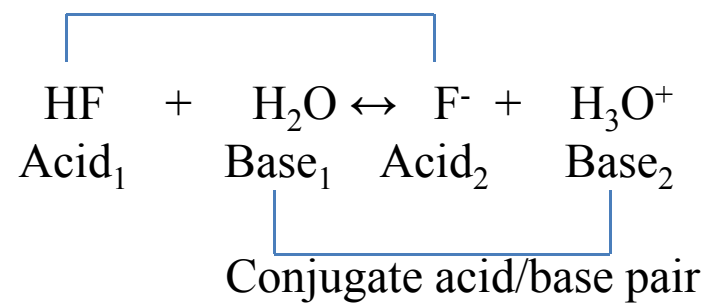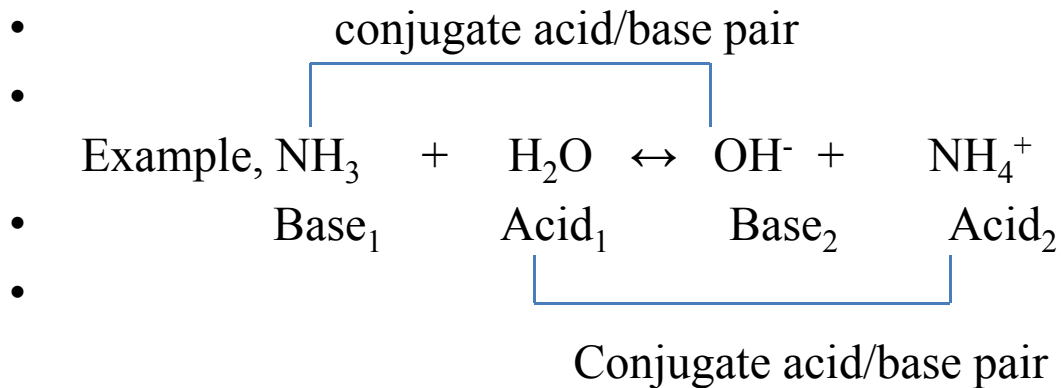-   Example I
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conjugate acid/base pair

$$HA \quad + \quad B \quad \leftrightarrow \quad A^- + \quad BH^+$$

$$Acid_1 \qquad Base_1 \quad Acid_2 \qquad Base_2$$

Conjugate base/acid pair
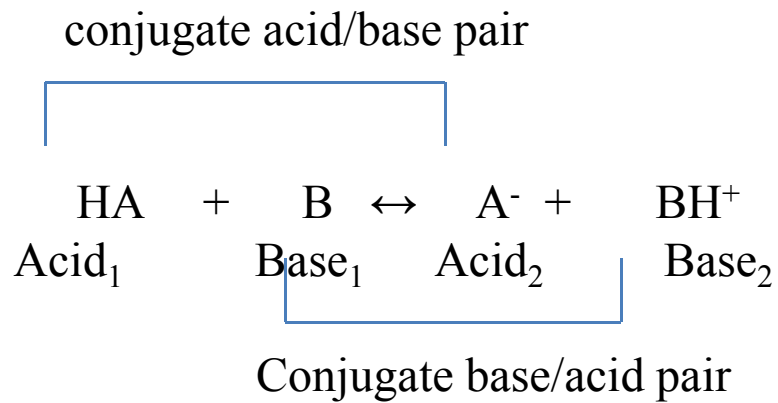
Conjugate acids/base pair

$$HF \quad + \quad H_2O \leftrightarrow F^- + \quad H_3O^+$$

$$Acid_1 \qquad Base_1 \quad Acid_2 \qquad Base_2$$

Conjugate acid/base pair

Module 11 Unit 2 cont'd

-
-          conjugate acid/base pair

Example, NH_3 + H_2O ↔ OH^- + NH_4^+

-         $Base_1$ $Acid_1$ $Base_2$ $Acid_2$

-

Conjugate acid/base pair

Strength of Acids and Bases

- Strength of acids and bases is the number of moles of H_3O^+/OH^- per mole of acid/base replicates

- Strong acids and bases dissociate completely in solution: $HCl ↔ H^+ + Cl^-$

- while weak acids and bases dissolves sparingly in solution:

- $CH_3OH ↔ CH_3O^- + H^+$

Module 11 Practice Question

1. (a) Define the following (i) Lewis acid (ii) BrØnstead-Lowry acid

2. List five properties of an acid?

3. Water from a river contaminated by alkali water has hydroxyl ion, $[OH^-]$, concentration of 3.7 x 10^{-3} mol/dm^3 , what is the pH of the river?

- Define Arrhenius or BrØnstead acids and bases?

- Define oxidation and reduction?

- Identify which of these two equations is an (i) oxidation and (ii) reduction?

- $2Mg \rightarrow 2Mg^{2+} + 4e^-$

- $O_2 + 4e^- \rightarrow 2O^{2-}$

- Which element in the above reaction is (i) a reducing agent (ii) an oxidizing agent?

- Identify the conjugate base or acid in the equation shown below?

$$NH_3 \quad + \quad H_2O \quad \leftrightarrow \quad OH^- \quad + \quad NH_4^+$$
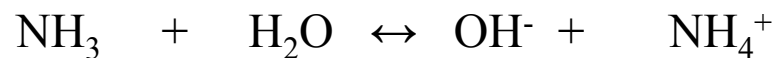
Module 12 Oxidation and Reduction

This chapter deals with the phenomenon of oxidation and reduction. Topics that will be discussed include:
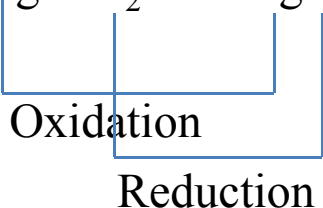
- Definition s of oxidation and reduction processes
- Differentiating between an oxidizing and reducing agent
- Half reactions
- Oxidation state or number
- Redox reactions
- Half Cells/Galvanic Cells: Daniel cells
- Gibbs Free Energy: Electrode potential
- Faradays Law of Electrolysis

Module 12 Unit 1
Oxidation and Reduction

Introduction

- Oxidation and reduction are complementary simultaneous reactions that occur naturally in nature. The abbreviation for the process is called REDOX (Reduction-Oxidation).

- Redox reaction is an equilibrium reaction. Examples of redox reaction include: (i), a piece of nail will rust within 2-3 days when exposed to the weather – $2Fe + 3/2O_2 \rightarrow Fe_2O_3$ (ii) Biological cells undergo oxidation (cancer cells) (iii) as part of the process of production of NaCl (Down's process) (iv) Galvanic cells: reaction that occurs in galvanic cells to produce electric current

- **Oxidation**: can be defined in 3 different ways: (1) loosing of hydrogen ions ($H_2CO_3 \rightarrow H^+ + HCO_3^-$) (ii) gaining of oxygen ($C + O_2 = CO_2$) (iii) loosing of electrons ($Cu \rightarrow Cu^{2+} + 2e^-$)

- **Reduction:** May be defined as (i) gaining of hydrogen ($H_2O + Cl_2 \rightarrow HCl + HOCl$) (ii) loosing of oxygen $2H_2O_2 \rightarrow 2H_2O + O_2$ (iii) $Cu^{2+} + 2e^- \rightarrow Cu$

- Other Examples: $2Mg + O_2 \rightarrow 2MgO$

-

- Oxidation

- Reduction

Module 12 Unit 1 cont'd

Half Reactions

- Involves expressing oxidation and reduction as separate equations
- $2Mg + O_2 \rightarrow 2MgO$ is equal to 2-half reactions:
- $2Mg \rightarrow 2Mg^{2+} + 4e^-$ -- Oxidation
- $O_2 + 4e^- \rightarrow 2O^{2-}$ -- Reduction
- **Reducing Agent**: Mg is losing electrons to reduce oxygen, hence it a reducing agent
- **Oxidizing Agent:** Oxygen is accepting electrons from Mg to oxidize Mg, hence it is an oxidizing agent
- **Oxidation State (Number)**
- Oxidation state is an ionic state achieved by an element (atom) either by gaining or losing electrons
- It is a convenient way of expressing a reaction process going on at a point in time
- The oxidation state of an element is zero irrespective of whether the element is an atom (Ne), a molecule (O_2, H_2) or a lattice (Si)

Module 12 Unit 1 cont'd

- In assigning an oxidation state (#), homonuclear bonds are ignored while heteronuclear bonds are considered. For example, In H_2O_2, the O-O bond is homonuclear and is not used in determining the oxidation state of oxygen in this compound. The heteronuclear bond, H-O, is considered in determining the oxidation state of H or O in this compound

- $2H \rightarrow H^+ + 2e^-$

- $O + e^{2-} \rightarrow O^{2-}$

- Oxidation process involves increase in oxidation number of an atom while reduction process involves decrease in oxidation number of an atom

- For example: $16HCl_l + 2KMnO_4(aq) \rightarrow 5Cl_{2(g)} + 2KCl_{(aq)} + 8H_2O_l + 2MnCl_{2(aq)}$

- Half Reactions for this reaction is as shown below

- $2Cl^- \rightarrow Cl_{2(g)} + 2e^-$

- Oxidation: -1   0

- Reduction: $Mn^{7+} + 5e^- \rightarrow Mn^{2+}$

- Reduction   +7   +2

- The net charge of ox/red process must be zero

Module 12 Unit 2

Half Cells/Galvanic Cells

- A simple electrochemical reaction cell consists of: (a) a metal strip dipping into a solution of one ion (Cu strip in Cu salt solution). There will be no chemical reaction in this set up.
- We may represent such cells as half cells: $Cu^{2+}(aq) + 2e^- \leftrightarrow Cu_s$
- We can put two such half cells together to form an electrical circuit.
- If there is a potential difference between the 2 half cells, a reaction will occur e.g. Daniel Cell
- Daniel Cell: consist of 2 half cells – Cu^{2+}/Cu and Zn^{2+}/Zn
- $Cu^{2+}_{aq} + 2e \rightarrow Cu_s$
- $Zn \rightarrow Zn^{2+} + 2e^-$
- When the two cells are connected by a salt bridge (gelatin impregnated with KCl or KNO_3), ions pass from one cell to another
- The two cells gradually mixes together while Redox reactions takes place
- $Zn_s + Cu^{2+}_{aq} \rightarrow Zn^{2+}_{aq} + Cu_s$
- This arrangement result into a Galvanic cell
- An electrical work is done in this type of cell which allows the mixing to occur
- The potential difference, E_{cell}, between the two half cells is measured in volts (V) with a voltmeter

Module 12 Unit 2 cont'd

- E_{cell} is related to change in Gibbs energy for the cell reaction and is also called the Standard cell potential under standard condition.

- $\Delta G_{(Gibbs\ energy)} = -ZFEo$

  Z = # of moles of e⁻'s transferred per mole of reaction. F = Faradays constant = 96485 C/mol, E^0_{Cell} = potential difference in Volts. ΔG = J/mol

salt bridge

Cu foil
electrode

Zn foil
electrode

Cu^{2+} salt

Zn^{2+} salt
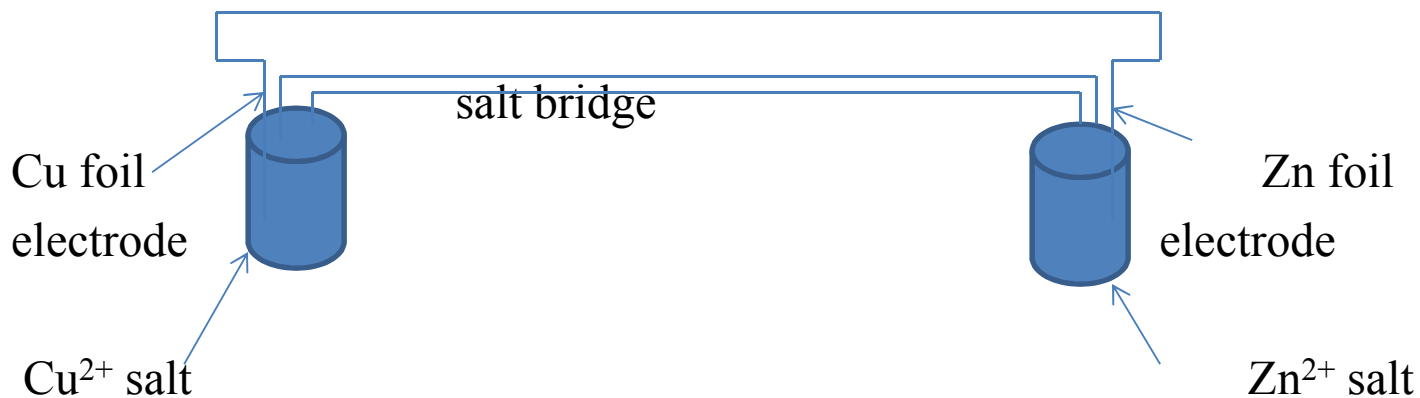
Module 12 Practice Questions

1. Describe a Daniel cell?
2. Describe a Galvanic cell?
3. What is the difference between a Daniel and Galvanic cell?
4. Draw a diagram that represent a typical Galvanic cell?
5. Explain the terms in the equation $\Delta G_{(Gibbs\ energy)}= -ZFEo$
6. What exactly does $\Delta G_{(Gibbs\ energy)}$ represent?
7. Define the oxidation and reduction process?
8. Define an oxidizing and reducing agent and give examples?
9. Redox reaction is an equilibrium process. Explain?
10. Give examples of redox processes?

Module 13 Unit 1

The topics in this module includes

- Electrolysis
- Electrodes and electrolytic cell
- Electrolysis of concentrated and dilute solutions
- Electroplating
- Faradays first and second law of Electrolysis

Module 13 Unit 1

Faradays Law of Electrolysis

- **First Law**
- The quantity of a substance produced by electrolysis is proportional to the quantity of electricity used.
- **Second Law**
- For a given quantity of electricity the quantity of substance produced is proportional to its weight.
- The quantity of electricity or charge contained in a current running for a specified time can be calculated:
- $Q = I \times t$
- Q = quantity of electricity or charge in coulombs (C)
  I = current in amps (A)
  t = time (seconds)
- The Faraday constant, F, is the quantity of electricity carried by one mole of electrons.
- F = Avogadro's $\times$ charge on electron in coulombs
- $F = 6.022 \times 10^{23} \text{ mol}^{-1} \times 1.602192 \times 10^{-19} \text{ C}$
  $F = 96,484 \text{ C mol}^{-1}$
  This is usually rounded off to 96,500 C mol^{-1} for calculations in chemistry.

Module 13 Unit 1 cont'd

- The quantity of electricity required to deposit an amount of metal can be calculated:
- $Q = n(e) \times F$
- Q = quantity of electricity in coulombs (C)
  $n(e)$ = moles of electrons
  F = Faraday constant = 96, 500 C mol^{-1}
- Electrical Energy, E, can be calculated:
- $E = Q \times V$
- E = electrical energy in joules (J)
  Q = quantity of electricity in coulombs (C)
  V = voltage (or EMF) in volts (V)
- 1 kilowatt-hour, kWH, is a unit of electrical energy.
- $1 \text{ kWH} = 3.6 \times 10^6 \text{ J}$
- **Examples**
- $Q = I \times t$
- Calculate the quantity of electricity, Q, obtained when a current of 25 amps runs for 1 minute.
- $Q = ? \text{ C}$
  $I = 25 \text{ A}$
  $t = 1 \text{ minute} = 60 \text{ seconds}$
- $Q = 25 \times 60 = 1,500 \text{ C}$
- .

Module 13 Unit 3

Example 2

- $I = Q \div t$
- Calculate the current needed to provide 30,000 coulombs of electricity in 5 minutes.
- $Q = 30,000$ C
  $I = ?$ A
  $t = 5$ minutes $= 5 \times 60 = 300$ seconds
- $I = Q \div t = 30,000 \div 300 = 100$ amps

Example 3

- $t = Q \div I$
- Calculate the time required to produce 12,000 C of electricity using a current of 10 amps.
  $Q = 12,000$ C
  $I = 10$ A
  $t = ?$
- $t = Q \div I = 12,000 \div 10 = 1,200$ seconds $= 1,200 \div 60 = 20$ minutes

Module 13 Unit 3 cont'd

Example 4

- $Q = n(e) \times F$

- Calculate the quantity of electricity obtained from 2 moles of electrons
  $Q = n \times F$
  $Q = ?$
  $n = 2$ mol
  $F = 96,500$ C mol^{-1}

- $Q = 2 \times 96,500 = 193,000$ C

Example 5

- $n(e) = Q \div F$

- Calculate the moles of electrons obtained from 250 C of electricity
  $n(e) = ?$ mol
  $Q = 250$ C
  $F = 96,500$ C mol^{-1}

- $n(e) = 250 \div 96,500 = 2.59 \times 10^{-3}$ mol

Module 13 Unit 3 cont'd

Example 6

- Calculate the time required to deposit 56g of silver from a silver nitrate solution using a current of 4.5A.

Calculate the moles of electrons required for the reaction:

- $Ag^+ + e \rightarrow Ag(s)$
  moles of Ag(s) deposited, n(Ag) = moles of electrons required, n(e)
  moles of Ag = n(Ag) = mass ÷ MM
  mass Ag deposited = 56g
  MM = 107.9 g mol^{-1} (from Periodic Table)
  n (Ag) 56 ÷ 107.9 = 0.519 mol = n(e)

- Calculate the quantity of electricity required: Q = n(e) x F
  Q = ? C
  n(e) = 0.519 mol
  F = 96,500 C mol^{-1}
  Q = 0.519 x 96,500 = 50,083.5 C

- Calculate the time required: t = Q ÷ I
  Q = 50,083.5 C
  I = 4.5 A
  t = 50,083.5 ÷ 4.5 = 11,129.67 seconds
  t = 11,129.67 ÷ 60 = 185.5 minutes
  t = 185.5 ÷ 60 = 3.1 hours

Example 7

- What mass of copper could be deposited from a copper (II) sulphate solution using a current of 0.50 A over 10 seconds?
    - Calculate the quantity of electricity: $Q = I \times t$
      $I = 0.50$ A
      $t = 10$ seconds
      $Q = 0.50 \times 10 = 5.0$ C
    - Calculate the moles of electrons: $n(e) = Q \div F$
      $Q = 5.0$ C
      $F = 96{,}500$ C mol^{-1}
      $n(e) = 5.0 \div 96{,}500 = 5.18 \times 10^{-5}$ mol
    - Calculate mass of copper: mass $= n \times MM$
      $Cu^{2+} + 2e \rightarrow Cu(s)$
      1 mole of copper is deposited from 2 moles electrons
      $n(Cu) = \frac{1}{2}n(e) = \frac{1}{2} \times 5.18 \times 10^{-5} = 2.59 \times 10^{-5}$ mol
      $MM = 63.55$ g mol^{-1} (from Periodic Table)
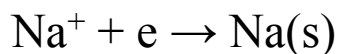   mass $(Cu) = (2.59 \times 10^{-5}) \times 63.55 = 1.65 \times 10^{-3}$ g $= 1.65$ mg

Module 13 Unit 3 cont'd

Example 8

- An EMF of 4.5 V produces 1 kg of sodium metal by the electrolysis of Na^+. Calculate the minimum number of kilowatt-hours of electricity needed to produce the sodium metal.

  - Calculate the moles of electrons, n(e), required
    Write the equation for the electrolysis of Na^+:
    $Na^+ + e \rightarrow Na(s)$
    moles of Na(s) = moles of electrons used n(e)
    n(e) = n(Na) = mass ÷ MM
    mass = 1kg = 1,000g
    MM = 22.99 g mol^{-1} (from Periodic Table)
    n(e) = 1,000 ÷ 22.99 = 43.497 mol

  - Calculate the quantity of electricity required: Q = n(e) x F
    n(e) = 43.497 mol
    F = 96,500 C mol^{-1}

    Q = n(e) x F = 43.497 x 96,500 = 4.2 x 10^6 C

Example 8 cont'd

- Calculate the electrical energy: $E = Q \times V$
  $Q = 4.2 \times 10^6$
  $V = 4.5$ V
  $E = 4.2 \times 10^6 \times 4.5 = 1.89 \times 10^7$ J

- Convert Electrical Energy to kilowatt-hours: kilowatt-hours $= E \div 3.6 \times 10^6$
  $E = (1.89 \times 10^7) \div (3.6 \times 10^6) = 5.25$ kWH

Module 13 Practice Questions

1. State Faradays first and second law?
2. Express Faradays first and second law in form of an equation?
3. Define Electroplating?
4. What is a cathode and an anode electrode?
5. What products will be produced if the concentrated an dilute solution of copper sulfate is subjected to electrolysis?

Module 14 Solutions

The topics in this module are:

- Solutions
- Properties of Solutions
- Types of solutions
- Water
- How a Solution is formed
- Electrolytes and Non Electrolytes
- Solubility
- Solubility
- Effect of Temperature on Solutions
- Henry's Law: Solubility vs. Pressure
- Soluble and Insoluble Salts
- %Composition, Molarity, dilution of a solution (dilution factor)
- Moles and Molarity
- Colligative Properties of a solution: Boiling and Freezing point, changes in Freezing and boiling points
- Molatity
- Osmotic Pressure, types of solutions

Solutions

Solutions

- Are everywhere around us
- Consist of a solute that is dissolve in a solvent e.g. coffee dissolved in hot water, sugar in milk or water, air = nitrogen + oxygen, salt + water, carbonated drinks (CO_2 + water + sugar), dissolved particles in the blood plasma etc
- Hence, we can define a solution as a mixture of 2 or more substances with no chemical bonding
- Properties of a Solution
- Components are physically separated
- Components behaves independently as if they were alone e.g. taste of salt in water, taste of coffee in a coffee drink etc
- Types of Solutions
- There are two types of solutions
- **Homogeneous:** components cannot be distinguished from each other e.g. syrup = sugar + water, salt + water. Components (solute) dissolve in each other (solvent). Solute is usually the smaller amount.

- **Heterogeneous:** components are distinguishable from each other e.g. an aquarium, gas in gas solution (oxygen + nitrogen), Gas in liquid solution (oxygen + water), Liquid + liquid = vinegar + water, solid + liquid (salt + water), liquid + solid (Hg + silver), solid + solid (brass= Zn + Cu)

Water

- Water is a universal solvent. It is made of 2 atoms of hydrogen and one of oxygen. Its shape is bent and is a covalent polar molecule making it possible for hydrogen bonding to be formed among its molecules

- Body of human adult is 60% water and that of an infant is 78%. Water in human cell is 60%. Extracellular fluid is 10% water. Loss of 10% of water leads to dehydration and a loss of >20% is very fatal and may lead to death

- Dehydration may be relieved by food and water

How a Solution is formed

- Solute and solvent particles must have similar polarity

- Both particles must attract each other

- Enough energy to break the bonds between particles of the solvent and solute

- After bond breaking, each particle is free to move and rearrange themselves

Module 14 Unit 2
Electrolytes, Non-Electrolytes and Solubility

Electrolytes/Non Electrolytes

- **Electrolytes:** are solutes that dissolve in solvents and can conduct electricity. Dissolves in water to form ions

- are classified according to the extent of dissolution in water: **Strong**: are completely dissolve in water ($NaCl + water = Na^+ + Cl^-$) **Weak:** partially dissolves in water, e.g. HF, CH_3COOH etc, forms small amount of ions in water e.g. $HF_{aq} \leftrightarrow H^+_{aq} + F^-_{aq}$

- **Non Electrolytes:** does not conduct electricity and does not dissolve in solvents to form ions. Dissolves in solvent as molecules e.g. $C_{12}H_{22}O_{11} \leftrightarrow C_{12}H_{22}O_{11aq}$

- Solubility

- Amount of solute that can dissolve in a given amount of solvent

- **Factors that affect Solubility:** Type of solvent, solute, concentration, and temperature

- **Units of solubility: gram of solute/100g of solvent**

- **Types of Saturation**: **Unsaturation**- solvent that can dissolve more solute; **Saturated**- solvent that has dissolved the maximum amount of solute it can dissolve. Any additional solute added to it will crystallize out of solution.
  Super Saturation: If a saturated solution is gradually cooled, it might be able to hold the solute in solution at this lower temperature than it should contain. If there is any agitation or any extra solute is added, crystallization will occur

Effect of Temperature on Saturation

- Increase in temperature increases solubility for some substances in solution e.g. sugar solubility increases in hot tea with increasing temperature. The converse is also true

- For a gas/water pair, solubility decreases with increase in temperature. For example, oxygen solubility in water decreases with increase in temperature, hence marine animals may die if water temperature increases.

Module 14 Unit 3
Henry's Law and Unit of Concentration

Henry's Law: Solubility vs. Pressure

- Henry's law states that solubility is directly related to pressure of the liquid. This implies that as pressure increases, solubility also increases. Example includes solubility of CO_2 in carbonated drinks. Carbonation is done at high pressure

- Soluble and Insoluble Salts

- Ionic salts can be soluble(NaCl) or insoluble in water ($CaCO_3$)

- Soluble salts usually contain the following: Li^+, Na^+, K^+, NH_4^+, NO_3^-, CH_3COO^-

- Some Cl^- salt may be soluble (NaCl) or insoluble (AgCl)

- % Composition

- Concentration (C): amount of solute dissolved in a certain amount of solution

- C = amt of solute/amount of solvent

- **Types of Concentration Unit:**

- (i) Mass % (m/m) = (mass of solute (g)/100g of solution) x 100%. It implies that Mass % = Mass of solute (g)/ mass of solute + mass of solvent

- Units of solute and solvent must be the same

Module 14 Unit 3 cont'd

- **Example**: (i) 15% (m/m) = 15 g of solute/100 g of solution   (ii) An antibiotic ointment is 3,5 % (m/m) of neomycin. How many grams of neomycin are in a tube containing 64 g of ointment?

- Answer: Mass % (m/m) = 3.5% i.e. 100g of ointment contain 3.5 g neomycin

- Therefore, 64 g of ointment will contain (3.5 x 64/100) of neomycin = 2.2 g

- (ii) Volume % (v/v) is used for gases and liquids

- V% (v/v) = (volume of solute/volume of solution) x 100% or (Volume of solute/ 100 mL of solution) x 100%

- Example: 15% (v/v) = 15 g of solute/100 mL of solution

- Molarity of Solution

- **Molarity (M)** = # of moles of solute/Liter of solution or   M = wt in grams/Molar mass x Liter of solution

- Example: How many liters of 2.0 M solution are needed to produce 67.5 g of NaCl?

- Answer: M = wt in grams/Molar mass x Liter of solution

- L = 67.5/2 x 55.8 = 0.576 L

Module 14 Unit 3 cont'd

Dilution:

- A lower concentrated solution may be prepared from a more concentrated one by adding more solvent (water). For example, adding 2 mL of water to a cup of juice
- # of moles = Molarity x Volume = M x V = constant on dilution
- The # of moles of the solute remains the same on dilution but the volume of solution changes, therefore,
- The # moles of solute in a concentrated solution = # moles of the solute in the diluted solution, i.e. $M_1V_1 = M_2V_2$
- Example: What volume of a 0.02M $SrCl_2$ can be prepared by diluting 50.0 mL of 1.0 M $SrCl_2$ solution?
- Answer: If $M_1V_1 = M_2V_2$.
- $\qquad$ $V_2 = M_1V_1/M_2$ = 50.0 x 1.0/0.02 = 2500 mL

Moles and Molarity

- $Zn(s) + 2HCl_{aq} \rightarrow ZnCl_{2aq}$
- How many liters of 1.5M HCl will react 5.3 g of Zn? Molar mass of Zn = 65.4 g
- Answer: # moles of Zn = mass of Zn/Molar mass = 5.3/65.4 = 0.081.
- From the equation 1 mole of Zn = 2 moles of HCl, therefore, moles of HCl = 2.0.081 = 0.162 moles
- Molarity of HCl = 1.5M, i.e. 1.5 moles = 1 liter
- Therefore, 0.162 moles = (1/1.5) x 0.162 L = 0.1 L

Properties of a Solution

- Dissolving a solute in a solvent will affect the properties of the solution
- There are two types of solutions: Homogeneous and non-homogeneous
- **Homogeneous Solution**: Solute particles are small and can pass through cells , semi-permeable membranes and filters e.g. NaCl solution. The solutes are equally mixed with the solvent everywhere in the bulk of the solution.
- Other examples include colloids, protein solutions, shaving cream, soap suds, butter, fog, cloud, dust etc
- **Colloids**: have large size particles but it is homogeneous. Can pass through filters but not semi-permeable membranes. Examples of colloids in the body are: starch and protein, fibers. These colloids cannot go through intestinal membranes except if it is digested and broken down to smaller molecules like glucose
- Colloids like fiber, are not digestible and hence are excreted out of the system
- **Non-Homogeneous**: non uniform mixtures
- **Suspensions:** are heterogeneous, non uniform mixtures that cannot pass through filters because they are large. Examples include, Gari, mud, kaopectate (Bismuth solution), calamine lotion etc. Often, they precipitate out of solution
- **Freezing Point(FP)/Boiling Point (BP):** Adding solute to a solution will change its physical properties such as freezing point (FP) or boiling point (BP)
- **FP and BP** are called **colligative** properties of a solvent
- **Example:** spreading urea salt on snow lowers it freezing point and prevent it from freezing. Adding antifreeze to a radiator prevents water from freezing the ethylene glycol in the radiator (decreases FP to 34^0C and raises the BP to 255^{0C})

Module 14 Unit 4 cont'd

- **Molality:** To calculate the change in FP or BP, we use the concentration unit called Molality = moles of solute/Kg of solvent. (MW glucose= 180.2 g/mole)

- Example: Calculate the molality of a solution containing 35.5g of glucose in 0.401 kg of water

- **Answer**: Molality= # moles of glucose/Kg of solvent

- M= 35.5/180.2 x 0.401 = 0.493M

Change in FP and BP

- The amount of solutes dissolved in a solvent will determine the extent of how FP will be depressed or BP will be elevated

- Increasing the amount of solute depresses FP but increases BP

- Depression in FP = ΔT_f and elevation of BP = ΔTb

- If 1 mole of a solute is dissolved in 1 kg of a solvent:

- ΔT_f= Molality x k_f (k_f = freezing point constant) and

- ΔTb= Molality x k_b ((k_f = boiling constant)

- For water, 1 Molal of water, k_f = 1.86^0C and hence , ΔT_f= 1.86^0C

Osmotic Pressure

Osmotic Pressure

- In biological processes, water and solvents moves in and out of cells. The amount of solute in solutions affect these movements

- In osmosis water molecules moves from area of low concentration of solute to area of higher concentration of solute

- If we put a semi-permeable membrane in the middle of a beaker and put water on one side and sugar solution on the other side. Water moves to the sugar solution side with time. Initially, the volume of water level decreases while that of the sugar solution increases. Later on, the volume on both sides will become the same. This process is called Osmosis

- Equilibration of volume on both sides is due to the pressure corresponding to the rise in height of the sugar solution (pd). This pressure is called the **osmotic pressure** of the sugar solution

- Increase in amount of solute increases the osmotic pressure

- **Reversed Osmosis**: involves applying external pressure to a substance so as to reverse osmosis. Water will move from area of higher concentration to area of lower concentration of the solute. Here, the applied pressure is greater than osmotic pressure. Example include the process of getting pure water from the salty ocean water (desalting process)

Module 14 Unit 5 cont'd

- **Isotonic Solution:** Fluids such as blood, tissue fluids, lymph, and plasma exert osmotic pressure. In-vitro (IV) solutions are isotonic (similar concentration like body fluids) and exert osmotic pressure like body fluids. IV fluids uses the units of mass/volume %. For example, 0.9% NaCl is the same as 0.9g of NaCl /100 mL of solvent

- If we place a red blood cell in an isotonic solution, it will retain its volume because its osmotic pressure is the same as that of the isotonic solution. The flow of water will be equal in both direction

Hypotonic/Hypertonic Solutions

- If a red blood cell is placed in a non isotonic solution, the solutions will affect the volume of the cell

- **Hypotonic Solution:** the osmotic pressure and the amount of solute is less than that of an isotonic solution and the blood cell. The amount of solute in the cell is higher than that of the hypotonic solution, hence, water will flow into the cell, its volume will increase and the cell will die. This process is called **hemolysis**

- **Hypertonic: has** more solute than an isotonic solution and blood cells. Water will flow from the cell to the container. The cell will be dehydrated and will die. The process is called **Crenation.**

Module 14 Practice Questions

- Define a hypotonic, isotonic and hypertonic solution?

- Name and describe the two known types of solution?

- Which of the following solvent is described as a universal solvent: sodium hydroxide, hydrochloric acid, water?

- Define and give examples of the following (i) strong electrolyte (ii) weak electrolyte (iii) and a non- electrolyte?

- Define (i) solubility and (ii) List the factors that affect solubility?

- List and describe the three types of solubility of a given solution?

- Describe how temperature, concentration, and pressure affect solubility?

- An antibiotic ointment is 3.5 % (m/m) of neomycin. How many grams of neomycin are in a tube containing 64 g of ointment?

- What volume of a 0.02M $SrCl_2$ can be prepared by diluting 50.0 mL of 1.0 M $SrCl_2$ solution?

- $Zn(s) + 2HCl_{aq} \rightarrow ZnCl_{2aq}$. How many liters of 1.5M HCl will react 5.3 g of Zn? Molar mass of Zn = 65.4 g

- Describe and give examples of a colloid and a suspension solution?

- List two colligative properties of a solution?

- (i) Define Molality (ii) Calculate the molality of a solution containing 35.5g of glucose in 0.401 kg of water