Chemistry
IX

Resource material for Ziauddin board

Prepared by: Ms. Saima Pervaiz Ahmed under the supervision of Madam Naheed Muneer Siddiqui

Benchmarks

Students will display a sense of curiosity and wonder about the natural world and demonstrate an increasing awareness that this has led to new developments in science and technology. They will learn from books and other sources of information and reconstruct previously learned knowledge.

Students will describe and explain common properties, forms, and interactions between matter and energy; their transformations and applications in biological, chemical and physical systems.

Students will demonstrate an understanding of the impact of science and technology on society and use science and technology to identify problems and creatively address them in their personal, social and professional lives. They will explain how scientists decide what constitutes scientific knowledge; how science is related to other ways of knowing; and how people have contributed to and influenced developments in science.

Content

Class IX	
Chapter 1	**Fundamentals of chemistry**
Chapter 2	**Structure of atom**
Chapter 3	**Periodic table and periodicity of properties**
Chapter 4	**Structure of molecule**
Chapter 5	**Physical states of matter**
Chapter 6	**Solutions**
Chapter 7	**Electrochemistry**
Chapter 8	**Chemical Reactivity**
Practicals = 12	

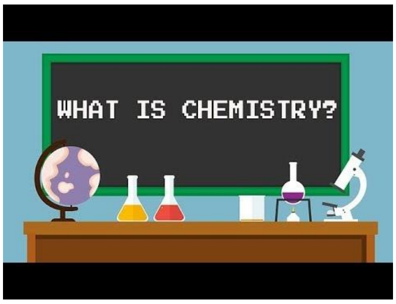
Chapter 01

Fundamentals of chemistry

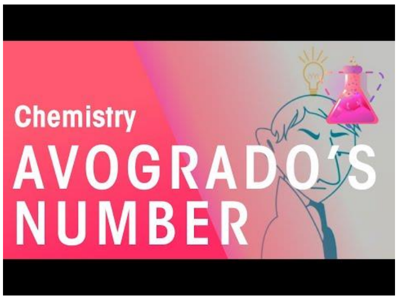
Chapter	Understanding	Skills	Practical
Fundamentals of chemistry **Topics according to national curriculum** • Branches of Chemistry • **Basic Definitions** Elements, Compounds and Mixtures ,Atomic Number, Mass Number • **Laws of chemical combination** • **Empirical Formula, Molecular Formula** Molecular Mass and Formula Mass • **Chemical Species** • **Ions (Cations, Anions) Avogadro's Number and Mole** • **Gram Atomic Mass, Gram Molecular and Gram Formula Mass Chemical Calculations** • **Mole-Mass Calculations** • **Balance chemical equation** • **Chemical reactions**	Student will: • Identify and provide examples of different branches of chemistry.(Applying) • Differentiate between branches of chemistry.(Understanding) • Distinguish between matter and a substance. (Analyzing) • Define ions, molecular ions, formula units and free radicals.(Remembering) • Define atomic number, atomic mass, atomic mass unit..(Remembering)	Students will be able to: • Calculate the number of representative particles in a given number of moles of any substance. • Calculate the mass of one mole of any substance. • Calculate the number of moles in a given	**Separate the given mixture by physical method**

	• Differentiate among elements, compounds and mixtures. (Remembering) • Differentiate between empirical and molecular formula. (understanding) • Distinguish among the term's gram atomic mass, gram molecular mass and gram formula mass and Avogadro's number. (Analyzing)	mass of a substance.	

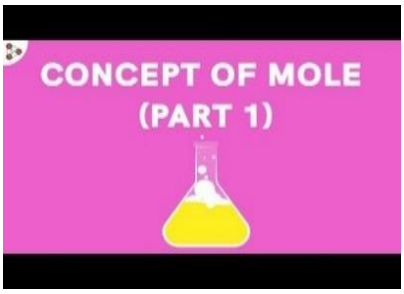
Videos



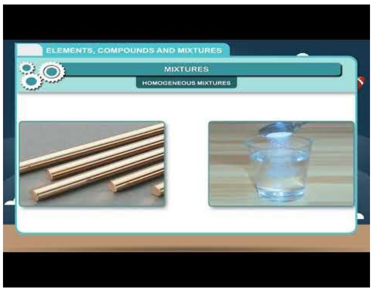
What is chemistry?



Avagadro's number

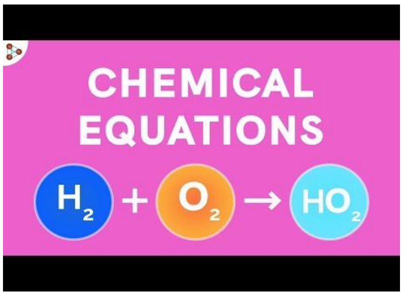


mole-I

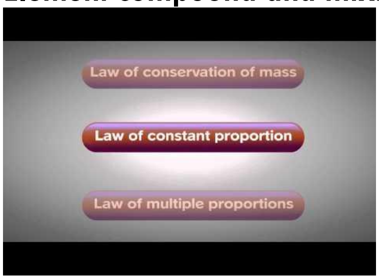


Element compound and mixture

mole-II



chemical equation



Laws of chemical combination

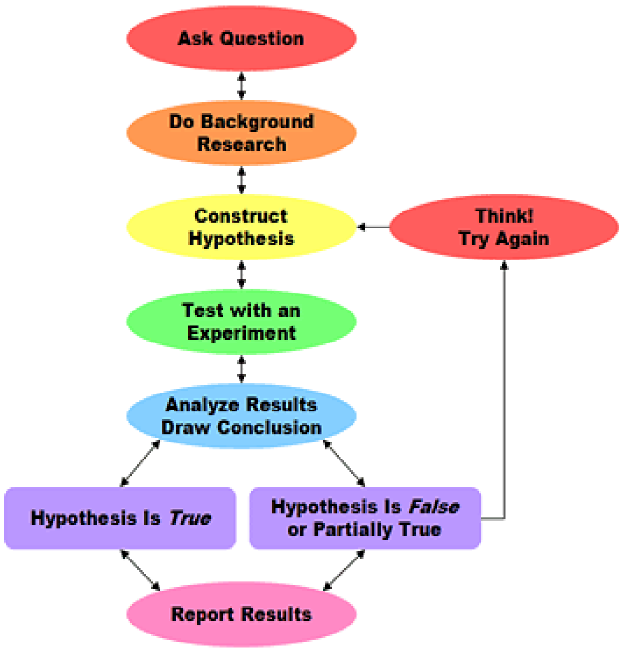
Chapter overview

What is chemistry

Chemistry is the study of matter, its properties, how and why substances combine or separate to form other substances, and how substances interact with energy.

Scientific approach in chemistry.

Scientific method way of discovering knowledge based on making falsifiable predictions (hypotheses), testing them, and developing theories based on collected data.



- In the scientific method, observations lead to questions that require answers.
- In the scientific method, the hypothesis is a testable statement proposed to answer a question.
- In the scientific method, experiments (often with controls and variables) are devised to test hypotheses.
- In the scientific method, analysis of the results of an experiment will lead to the hypothesis being accepted or rejected.

Branches of chemistry

1.Physical chemistry:

It is the branch of chemistry that deals with the laws and the principles governing the combination of atoms and molecules in chemical reactions. Physical chemists typically study the rate of a chemical reaction.

2. Organic chemistry:

Its involves the study of the structure, properties, and preparation of chemical compounds that consist primarily of carbon and hydrogen with the exception of CO2, CO, metal carbonates, bicarbonates

3. Inorganic chemistry:

Its deals with the chemistry of elements and their compounds, Inorganic chemists study things such as crystal structures, minerals, metals, catalysts, and most elements in the Periodic Table.

4. Analytical chemistry:

Its deals with the study of the methods and techniques. Analytical chemistry involves the qualitative and quantitative determination of the chemical components of substances

5. Bio-chemistry:

it is the study of chemical reactions that take place in living things. It tries to explain them in chemical terms.

6. Industrial and Applied chemistry:

Its deals with the study of different with physical and chemical processes towards the transformation of raw materials into products that are of beneficial to humanity.
Synthetic products like glass, cement, paper, soda and fertilizers etc.

7. Nuclear chemistry:

Its deals with the study of changes occurring in the nuclei of atoms, accompanied by the emission of invisible radiations.it is dealing with radioactivity, nuclear processes, and transformations in the nuclei of atoms

8. Environmental chemistry:

Its deals with the study of the interaction of materials and their effect on the environment of animals and plants. Personal hygiene, pollution and health hazards are important areas of environmental chemistry.

9. Polymeric chemistry:

It deals specially with the study of polymerization and the products obtained through the process of polymerization such as plastics, synthetic fibers, papers etc.

Basic definitions of Chemistry

Atom

The smallest particle of an element.
Scientists originally thought that atoms could not be split up, but this is not the case.

Molecule

A cluster of non-metal atoms that are chemically bonded together.
The atoms in a molecule are joined by covalent bonds.

Mixtures

A mixture is made from different substances that are not chemically joined.

Substance

A sample of matter that has the same physical and chemical properties throughout is called a substance. Chemistry recognizes two different types of substances: elements and compounds.

Element

An element is the simplest type of chemical substance; it cannot be broken down into simpler chemical substances by ordinary chemical means. There are 118 elements known to science, of which 80 are stable.

Compound

A compound is a combination of more than one element. There are over 50 million compounds known, and more are being discovered daily. Examples of compounds include water, penicillin, and sodium chloride (the chemical name for common table salt).

Mixtures

Physical combinations of more than one substance are called mixtures. There are two types of mixtures.

Heterogeneous mixture

A heterogeneous mixture is a mixture composed of two or more substances.

Homogeneous mixture/ Solution

An homogeneous mixture is a combination of two or more substances that is so intimately mixed that the mixture behaves as a single substance.

Matter – Anything that has mass and occupies space (volume).

Atomic number

The number of protons in an atom is called the atomic number.

Atomic mass or mass number

The number of protons and the number of neutrons determine an element's mass number: mass number = protons + neutrons.

Avogadro's number:

The number of atoms present in 12 g of carbon-12, which is 6.022×10 and the number of elementary entities (atoms or molecules) comprising one mole of a given substance.

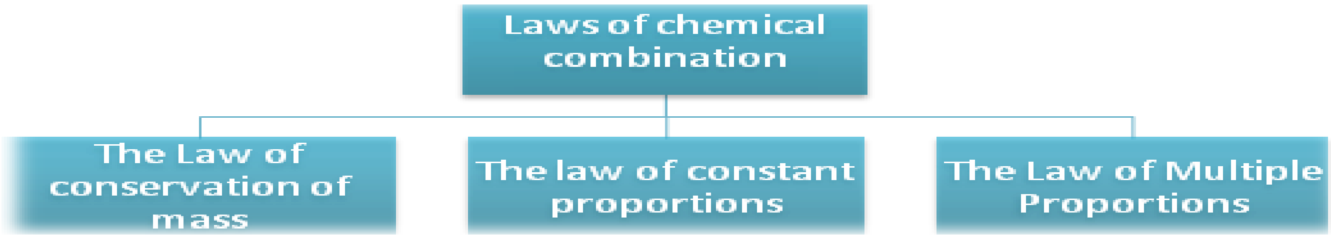
Empirical formulas show the simplest whole-number ratio of atoms in a compound. The empirical formula of glucose would be CH2O.

Molecular formulas show the number of each type of atom in a molecule. The molecular formula of glucose is C6H12O6,

The formula mass (formula weight) of a molecule is the sum of the atomic weights of the atoms in its empirical formula.

The molecular mass (molecular weight) of a molecule is its average mass as calculated by adding together the atomic weights of the atoms in the molecular formula.

Laws of chemical combination



1.Law of Conservation of Mass

In simple terms, this law states that matter can neither be created nor destroyed. In other words, the total mass, that is, the sum of mass of reacting mixture and the products formed remains constant. Antoine Lavoisier gave this law in the year 1789 based on the data he obtained after carefully studying numerous combustion reactions.

2. Law of Definite Proportions

Joseph Proust, a French chemist stated that the proportion of elements by weight in a given compound will always remain exactly the same. In simple terms we can say that, irrespective of its source, origin or its quantity, the percent composition of elements by weight in a given compound will always remain the same.

3. Law of Multiple Proportions

This law states that if two elements combine to form more than one compound, the masses of these elements in the reaction are in the ratio of small whole numbers. This law was given by Dalton in the year 1803.

Mole mass calculation.

Calculation of mole

Mole = given mass/atomic mass

Example:

number of moles = mass ÷ relative formula mass

Question: Calculate the number of moles of carbon dioxide molecules in 22 g of CO_2.

atomic mass of C = 12,

atomic mass of O = 16

formula mas of carbon dioxide = 12 + 16 + 16 = 44 (2 atoms of oxygen are present)

number of moles = 22 ÷ 44 = 0.5 mol

Exercise :

Calculate the mass of 2 mol of carbon dioxide (CO_2).

10 mole of carbon dioxide has a mass of 440 g. What is the relative formula mass of carbon dioxide?

Calculation of mass

Molecular Formula from Empirical Formula

the molecular formula of glucose is $C_6H_{12}O_6$ and empirical formula is $(CH_2O)_6$

The molar mass of the empirical formula CH_2O is 12 g C + 2(1 g H) + 16 g O = 30 g/mol.

The molar mass of glucose is 180 g/mol,

Molar Mass Calculation

Question

We begin by finding the atomic mass of each element in the periodic table.

The molar mass equals the sum of the atomic masses expressed in g/mol.

(a) The atomic mass of Na is 23 amu, and the molar mass of Sodium equals 23 g/mol.

(b) The sum of the atomic masses for NH_3 is 14.01 amu + 3(1.01)amu = 17.04 amu. The molar mass of ammonia equals 17.04 g/mol.

(c) The sum of the atomic masses for $Mg(NO_3)_2$ is 24.31 amu + 2(14.01 + 16.00 + 16.00 + 16.00)amu = 148.33 amu. The molar mass of magnesium nitrate equals 148.33 g/mol. Solution

Exercise for practice

Calculate the molar mass for each of the following substances: (a) manganese metal, Mn (b) sulfur hexafluoride, SF_6 (c) strontium acetate, $Sr(C_2H_3O_2)_2$ Answers: (a) 54.94 g/mol; (b) 146.07 g/mol; (c) 205.72 g/

Calculation of number of Atom and molecules (Avogadro's number)

Mass→Moles and Moles→Atoms

Example:

How many atoms of gold are in 58.27 g of gold?

The periodic table shows us that gold, Au, has the atomic weight 196.967 u. This means that its molar mass is 196.967 g/mol.

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First Step: Mass→Moles

To calculate moles of Au, divide the given mass by the molar mass.

58.27g Au / Au196.967g Au=0.295836 mol Au
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Second Step: Moles→Atoms

To calculate atoms of Au, multiply moles Au by Avogadro's number.

0.295836mol Au×6.022×1023atoms=1.782×1023atoms Au rounded to four significant figures due to 58.72 g.

Balancing Chemical Equations

Balancing chemical equations involves the addition of stoichiometric coefficients to the reactants and products. This is important because a chemical equation must obey the law of conservation of mass and the law of constant proportions, i.e. the same number of atoms of each element must exist on the reactant side and the product side of the equation.

Two quick and easy methods of balancing a chemical equation are discussed in this article. The first method is the traditional balancing method and the second one is the algebraic balancing method.

Balancing Method

The first step that must be followed while balancing chemical equations is to obtain the complete unbalanced equation. In order to illustrate this method, the combustion reaction between propane and oxygen is taken as an example.

Step 1

The unbalanced equation must be obtained from the chemical formulae of the reactants and the products (if it is not already provided).

The chemical formula of propane is C_3H_8. It burns with oxygen (O_2) to form carbon dioxide (CO_2) and water (H_2O)

The unbalanced chemical equation can be written as **$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$**

Step 2

The total number of atoms of each element on the reactant side and the product side must be compared. For this example, the number of atoms on each side can be tabulated as follows.

Chemical Equation: $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$	
Reactant Side	**Product Side**
3 Carbon atoms from C_3H_8	1 Carbon atom from CO_2
8 Hydrogen atoms from C_3H_8	2 Hydrogen atoms from H_2O
2 Oxygen atoms from O_2	3 Oxygen atoms, 2 from CO_2 and 1 from H_2O

Step 3

Now, stoichiometric coefficients are added to molecules containing an element which has a different number of atoms in the reactant side and the product side.

The coefficient must balance the number of atoms on each side.

Chemical Equation: $C_3H_8 + O_2 \rightarrow 3CO_2 + H_2O$	
Reactant Side	**Product Side**
3 Carbon atoms from C_3H_8	3 Carbon atoms from CO_2
8 Hydrogen atoms from C_3H_8	2 Hydrogen atoms from H_2O
2 Oxygen atoms from O_2	7 Oxygen atoms, 6 from CO_2 and 1 from H_2O

Step 4

Step 3 is repeated until all the number of atoms of the reacting elements are equal on the reactant and product side.

Chemical Equation: $C_3H_8 + O_2 \rightarrow 3CO_2 + 4H_2O$	
Reactant Side	**Product Side**
3 Carbon atoms from C_3H_8	3 Carbon atoms from CO_2
8 Hydrogen atoms from C_3H_8	8 Hydrogen atoms from H_2O
2 Oxygen atoms from O_2	10 Oxygen atoms, 6 from CO_2 and 4 from H_2O

Now that the hydrogen atoms are balanced, the next element to be balanced is oxygen.

Chemical Equation: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$	
Reactant Side	**Product Side**
3 Carbon atoms from C_3H_8	3 Carbon atoms from CO_2
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8 Hydrogen atoms from C_3H_8	8 Hydrogen atoms from H_2O
10 Oxygen atoms from O_2	10 Oxygen atoms, 6 from CO_2 and 4 from H_2O

Step 5

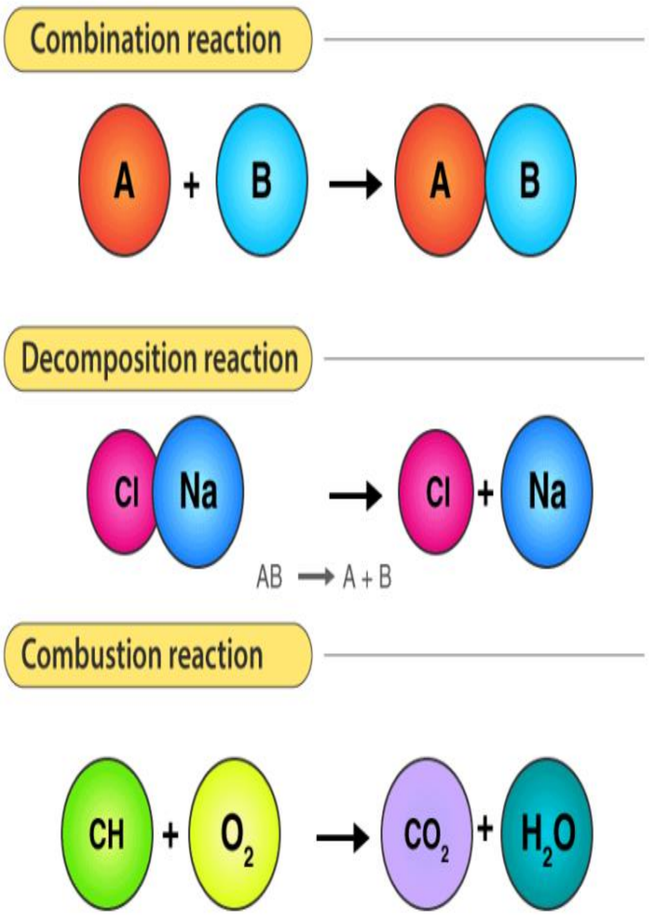
Once all the individual elements are balanced, the total number of atoms of each element on the reactant and product side are compared once again.
If there are no inequalities, the chemical equation is said to be balanced.
In this example, every element now has an equal number of atoms in the reactant and product side.
Therefore, the balanced chemical equation is **$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$.**

Different Types of Chemical Reactions

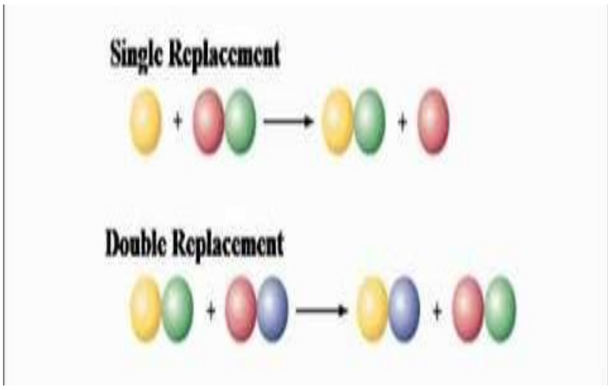
The types of chemical reaction are:
Combination reaction
Decomposition reaction
Displacement reaction
Double Displacement reaction
Precipitation Reaction

1. Combination Reaction

A reaction in which ***two or more reactants combine to form a single product*** is known as a combination reaction.
It takes the form of $X + Y \rightarrow XY$
Combination reaction is also known as a synthesis reaction.



Example of combination reaction: $2Na + Cl_2 \rightarrow 2NaCl$



2. Decomposition Reaction

A reaction in which a ***single compound breaks into two or more simpler compounds*** is known as a decomposition reaction.
It takes the form of $XY \rightarrow X + Y$
A decomposition reaction is just the opposite of combination reaction.
Example of a decomposition reaction: $CaCO_3 \rightarrow CaO + CO_2$
The reaction in which a compound decomposes due to heating is known as a thermal decomposition reaction.

3. Displacement Reaction

A chemical reaction in which a ***more reactive element displaces a less reactive element from its aqueous salt solution.***
It takes the form $X + YZ \rightarrow XZ + Y$
It is also called a substitution reaction
Example of displacement reaction: $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

4. Double Displacement Reaction

A chemical reaction in which ***ions gets exchanged between two reactants which form a new compound*** is called a double displacement reaction.
It takes the form of XY + ZA → XZ + YA
It is also called a metathesis reaction
Example of double displacement reaction:

$BBaCl_2 + Na_2SO_4 \rightarrow BaSO_4 + 2NaCl$

Reference pages

https://www.thoughtco.com/formula-mass-versus-molecular-mass-3976099
https://schoolworkhelper.net/general-chemistry-key-terms-definitions/
https://www.omicsonline.org/industrial-chemistry.php
https://chem.libretexts.org/Bookshelves/Introductory_Chemistry/Book%3A_Introductory_Chemistry_(CK-12)/01%3A_Introduction_to_Chemistry/1.03%3A_Areas_of_Chemistry
https://socratic.org/questions/what-are-the-branches-of-chemistry-and-their-definition
https://chem.libretexts.org/Bookshelves/Introductory_Chemistry/Book%3A_Beginning_Chemistry_(Ball)/01%3A_What_Is_Chemistry%3F/1.1%3A_Basic_Definitions
https://www.bbc.co.uk/bitesize/guides/z84wfrd/revision/3
https://socratic.org/questions/5903ea887c01493de9f14f4c
https://byjus.com/chemistry/laws-of-chemical-combination-for-elements-and-compounds/

Lesson plan

https://study.com/academy/lesson/what-is-chemistry-lesson-plan.html

worksheets

Molar Mass and Mole Calculations Worksheet

1. Calculate the molar mass for each of the following compounds:

a. $NaCl$
b. Fe_2O_3
c. $MgCO_3$
d. $Ba(OH)_2$
e. $(NH_4)_3PO_4$
f. $Ca_3(PO_4)_2$
g. $C_6H_{12}O_6$
h. C_5H_5N

2. Calculate the number of moles in each of the following masses:

a. 45.0 g of acetic acid, CH_3COOH

b. 7.04 g of lead (II) nitrate

3. Calculate the mass in grams of each of the following amounts:

a. 3.00 mol of selenium oxybromide, $SeOBr_2$

b. 488 mol of calcium carbonate

4. Calculate the number of molecules or formula units in each of the following amounts:

a. 4.27 mol of tungsten(VI) oxide, WO_3

b. 0.989 mol of nitrobenzene $C_6H_5NO_2$

5. Calculate the mass of each of the following quantities:

a. 8.39×10^{23} molecules of fluorine

Name: ______________________________ **Date:** __________

Balancing Chemical Equations

Balance the following chemical equations.

1. ____ Fe + ____ H_2S0_4 $\longrightarrow$ ____ $Fe_2(SO_4)_3$ + ____ H_2

2. ____ CH_4 + ____ O_2 $\longrightarrow$ ____ CO_2 + ____ H_2O

3. ____ $SiCl_4(\ell)$ + ____ $H_2O(\ell)$ $\longrightarrow$ ____ $SiO_2(s)$ + ____ $HCl(aq)$

4. ____ AgI + ____ Na_2S $\longrightarrow$ ____ Ag_2S + ____ NaI

5. ____ NH_3 + ____ O_2 $\longrightarrow$ ____ NO + ____ H_2O

6. ____ $FeO_3(s)$ + ____ $CO(g)$ $\longrightarrow$ ____ $Fe(\ell)$ + ____ $CO_2(g)$

7. ____ SiO_2 + ____ HF $\longrightarrow$ ____ SiF_4 + ____ H_2O

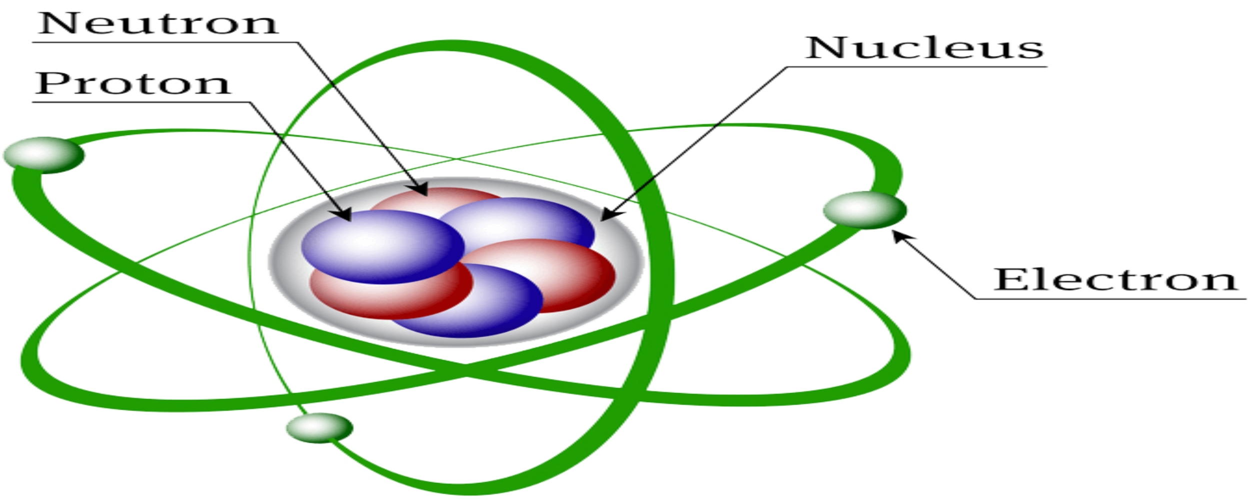
8. ____ $NaBr$ + ____ Cl_2 $\longrightarrow$ ____ $NaCl$ + ____ Br_2

9. ____ $(NH_4)_3PO_4$ + ____ $Pb(NO_3)_4$ $\longrightarrow$ ____ $Pb_3(PO_4)_4$ + ____ NH_4NO_3

10. ____ $Mg(OH)_2$ + ____ HCl $\longrightarrow$ ____ $MgCl_2$ + ____ H_2O

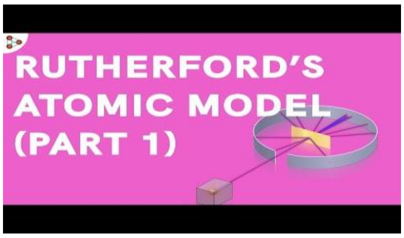
Chapter 02

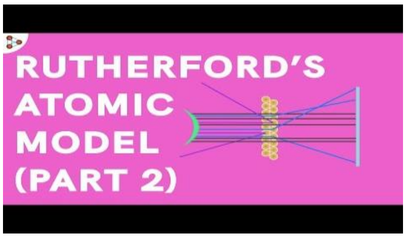
Structure of ATOM



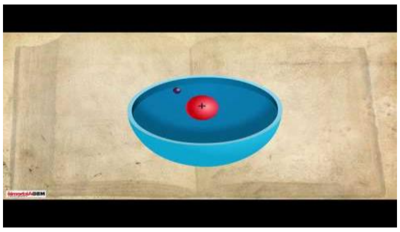
Chapter	Understandings
Structure of atom Topics according to national curriculum • **introduction** • **Discovery of sub atomic particle** • **Theories and experiments related to atomic structure.** Rutherford's atomic model Bohr's atomic model • Electronic configuration • Isotopes • Uses in daily life	Students will be able to: • Describe the contributions that Rutherford made to the development of the atomic theory. (Applying) • Explain how Bohr's atomic theory differed from its. (Analyzing) • Describe the structure of an atom including the location of the proton, electron and neutron. Define isotopes. (Remembering) • Compare isotopes of an Discuss properties of the isotopes of H, C, Cl, U. (Understanding) • Draw the structure of different isotopes from mass number and atomic number. (Applying) • Write the electronic configurations of the first 18 elements in the Periodic Table. (Remembering)

Videos

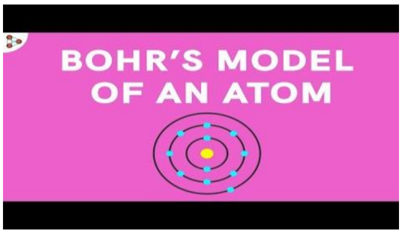




Rutherford's atomic model part 1 and 2



Bohar's atomic model



Bohar's atomic model

Electronic configration

Chapter overview

Discovery of atom

- The atom was first conceived of by the Greek philosopher Democritus in approximately 400 BCE.
- The concept was lost during the Dark Ages of Europe until 1803, when the British scientist John Dalton speculated that everything was composed of very tiny indivisible particles called atoms.
- Dalton's simple model of an atom persisted until 1897.
- When another British physicist, J.J. Thomson, discovered that atoms contained tiny negatively charged particles called electrons.
- From 1897 to 1909, scientists thought that atoms were composed of electrons spread uniformly throughout a positively charged matrix. J.J. Thomson's model was known as the plum pudding model.

Dalton's Atomic Theory

The English chemist **John Dalton** suggested that all matter is made up of **atoms,** which were indivisible and indestructible. He also stated that all the atoms of an element were exactly the same, but the atoms of different elements differ in size and mass.

The following are the postulates of his theory:

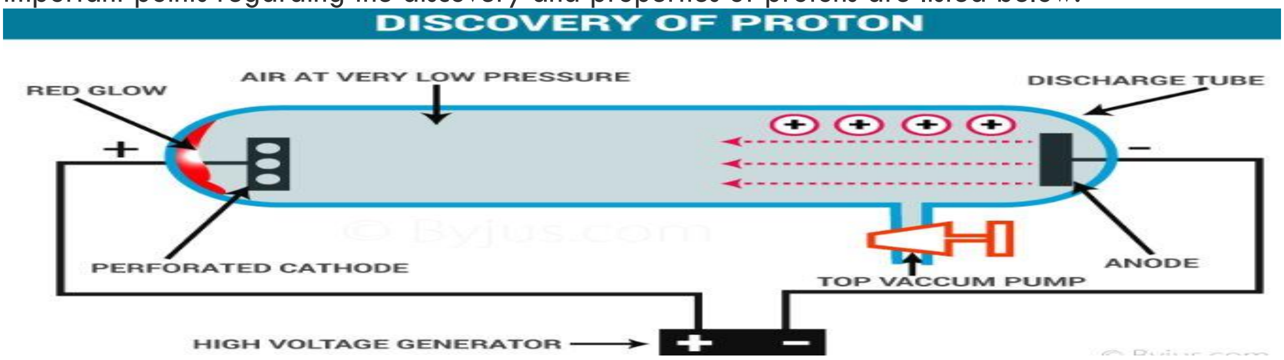
Every matter is made up of atoms.
Atoms are indivisible.
Specific elements have only one type of atoms in them.
Each atom has its own constant mass that varies from element to element.
Atoms undergo rearrangement during a chemical reaction.
Atoms can neither be created nor be destroyed but can be transformed from one form to another.

Discovery and Features of Subatomic Particles

The discovery of the three basic subatomic particles and some of their important features are discussed in this subsection.

Protons

Protons and Neutrons together make up the nucleus of an atom and are hence called nucleons. Some important points regarding the discovery and properties of protons are listed below.



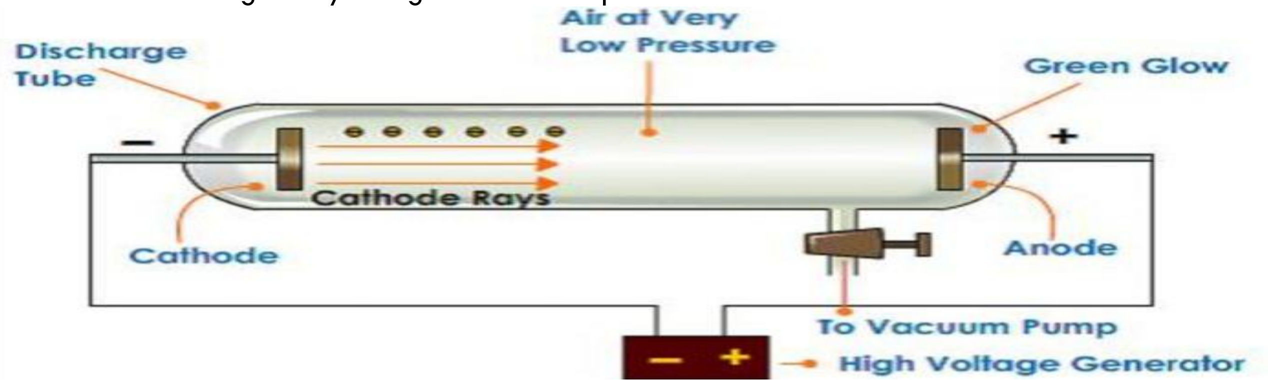
Protons are positively charged subatomic particles.

- The number of protons in an atom is equal to the number of electrons in it.
- The discovery of protons is credited to Ernest Rutherford.
- Protons can be produced via the removal of an electron from a hydrogen atom.
- The mass of a proton is $1.676 * 10^{-24}$ grams.
- The charge of a proton is $+1.602 * 10^{-19}$ Coulombs.

Electrons

Electrons are the subatomic particles that revolve around the nucleus of an atom. These electrons may be removed from or gained by an atom to form ions.

- Electrons are negatively charged subatomic particles



- An equal number of electrons and protons are found in the atoms of all elements.
- J. Thompson is credited with the discovery of electrons since he was the first person to accurately calculate the mass and the charge on an electron.
- The mass of an electron is negligible when compared to the mass of a proton. It is found to have a mass equal to (1/1837) times the mass of a proton.
- The charge of an electron is equal to -1.602 * 10^{-19} Coulombs.

Neutrons

Neutrons, along with protons, make up the nucleons. Neutrons are named for their neutral nature – unlike protons and electrons, they do not carry any charge. The discovery and general properties of neutrons are discussed below.

Neutrons are neutrally charged subatomic particles.

The masses of two different isotopes of an element vary due to the difference in the number of neutrons in their respective nuclei.

The neutron was discovered by James Chadwick in the year 1932.

They were discovered in an experiment wherein a thin sheet of beryllium was bombarded with alpha particles.

The mass of a neutron is 1.676 * 10^{-24} grams

Rutherford Atomic Model

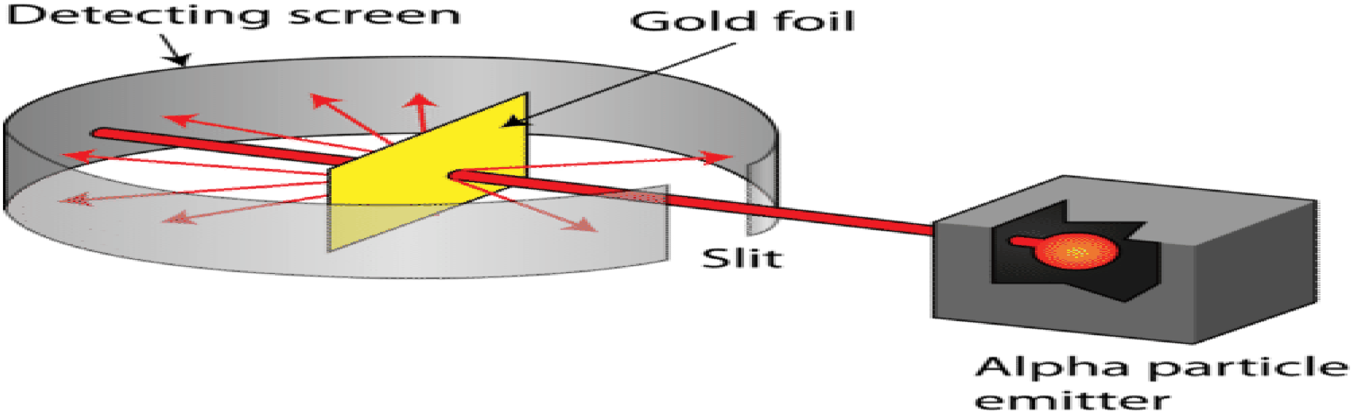
The plum pudding model is given by J. J. Thomson failed to explain certain experimental results associated with the atomic structure of elements.

Ernest Rutherford, a British scientist conducted an experiment and based on the observations of this experiment he proposed the atomic structure of elements and gave Rutherford Atomic Model.

Rutherford's Alpha Scattering Experiment

This model of an atom was developed by Ernest Rutherford, a New Zealand native working at the University of Manchester in England in the early 1900s.

- Rutherford spent most of his academic career researching aspects of radioactivity and, in 1908, won the Nobel Prize for his discoveries related to radioactivity. It was after this that Rutherford began developing his model of the atom
- Rutherford's conducted an experiment by bombarding a thin sheet of gold with α-particles and then studied the trajectory of these particles after their interaction with the gold foil.



Observations of Rutherford's Alpha Scattering Experiment

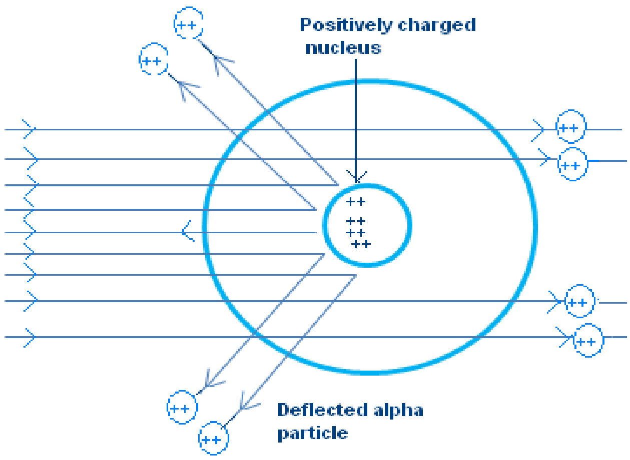
The observations made by Rutherford led him to conclude that:

- Most of the alpha particles bombarded towards the gold sheet passed straight through the gold foil. Means most of the space in an atom is empty.
- There was a deflection of the alpha particles by a small angle, and hence the positive charge in an atom is not uniformly distributed.
- Very few of the α-particles were deflected back, that is only a few α-particles had nearly 180° angle of deflection. So the volume occupied by the positively charged particles in an atom is very small as compared to the total volume of an atom.

Rutherford Atomic Model

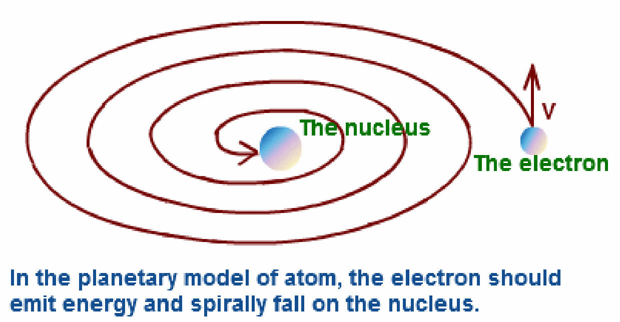
Based on the above observations and conclusions, Rutherford proposed the atomic structure of elements. According to the Rutherford atomic model:

- The positively charged particles of an atom was concentrated in an extremely small volume. He called this region of the atom as a nucleus.
- Rutherford model proposed that the negatively charged electrons surround the nucleus of an atom. He also claimed that the electrons surrounding the nucleus revolve around it with very high speed in circular paths. He named these circular paths as orbits.
- Electrons being negatively charged and nucleus being a densely concentrated mass of positively charged particles are held together by a strong electrostatic force of attraction.



What were the drawbacks of Rutherford's model of the atom?

- The revolution of the electron in a circular orbit is not expected to be stable.
- During acceleration, charged particles would radiate energy. Thus, the revolving electron would lose energy and finally fall into the nucleus.
- If this were so, the atom should be highly unstable and hence matter would not exist in the form that we know. We know that atoms are quite stable



Subatomic Particles of atom

Protons

Protons are positively charged subatomic particles. The charge of a proton is 1e, which corresponds to approximately 1.602×10^{-19}
The mass of a proton is approximately 1.672×10^{-24}
Protons are over 1800 times heavier than electrons.
The total number of protons in the atoms of an element is always equal to the atomic number of the element.

Neutrons

The mass of a neutron is almost the same as that of a proton i.e. 1.674×10^{-24}
Neutrons are electrically neutral particles and carry no charge.
Different isotopes of an element have the same number of protons but vary in the number of neutrons present in their respective nuclei.

Electrons

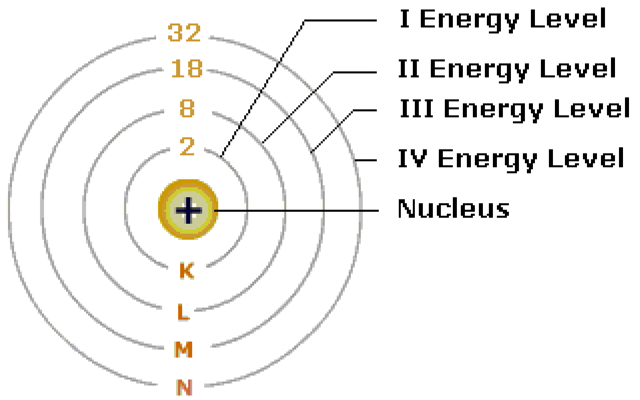
The charge of an electron is -1e, which approximates to -1.602×10^{-19}
The mass of an electron is approximately 9.1×10^{-31}.
Due to the relatively negligible mass of electrons, they are ignored when calculating the mass of an atom.

Bohr's Model of Atom

Neils Bohr, a Danish physicist, in 1913 proposed model of atom which rectified the problems left by Rutherford's Model. He proposed that

- Electrons revolve round the nucleus in a fixed orbit.
- He called these orbits as 'stationary orbit'.
- Each stationary orbit is associated with fixed amount of energy, thus electrons do not radiate energy as long as they keep on revolving around the nucleus in fixed orbit.
-

The circular path around the nucleus is called orbit, energy level or shell. Energy level are represented by letter – K, L, M, N, and so on.



Therefore,
1^{st} orbit is denoted by – K
2^{nd} orbit is denoted by – L
3^{rd} orbit is denoted by – M, and so on.
The orbits are denoted by 1, 2, 3, and so on.
Structure of The Atom
Distribution of Electrons in a Orbit or Shell:
The distribution of electrons in an orbit can be obtained by using formulae $2n^2$
Where 'n' is number of orbit.

Number of electrons in K-shell i.e. in 1^{st} orbit.

Here n = 1
Therefore,

$2n^2 = 2(1)^2 = 2$

Number of electrons in L-shell, i.e. in 2nd orbit

$2n^2 = 2(2)^2 = 8$

Number of electrons in M-shell, i.e. in 3rd orbit

$2n^2 = 2(3)^2 = 18$

In similar way maximum number of electrons in any shell can be calculated.

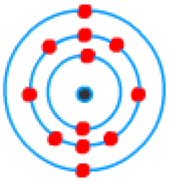
Introduction to what are Electron Configurations?

The electron configuration of an element describes how electrons are distributed in its atomic orbitals.

Electronic Configuration of Elements - From Magnesium to Calcium

Electronic configuration of Magnesium

Atomic number of magnesium = 12



Magnesium

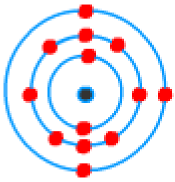
Therefore number of electrons = 1

Thus, electronic configuration of magnesium is $\begin{matrix} K & L & M \\ 2 & 8 & 2 \end{matrix}$

Number of orbit in magnesium = 3.

Electronic configuration of Aluminium

Atomic number of aluminium = 13.



Aluminium

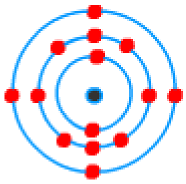
Therefore number of electrons = 13.

Thus, electronic configuration of aluminium is $\begin{matrix} K & L & M \\ 2 & 8 & 3 \end{matrix}$

Number of orbit in aluminium = 3

Electronic configuration of Silicon

Atomic number of silicon = 14



Silicon

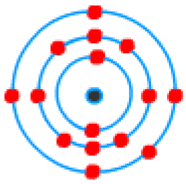
Therefore number of electrons = 14

Thus, electronic configuration of silicon is $\begin{matrix} K & L & M \\ 2 & 8 & 4 \end{matrix}$

Number of orbit in silicon = 3

Electronic configuration of Phosphorous (P)

Atomic number of phosphorous = 15



Phosphorous

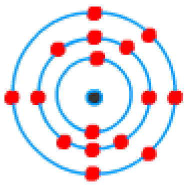
Therefore, number of electrons = 15

Thus, electronic configuration of phosphorous is $\begin{matrix} K & L & M \\ 2 & 8 & 5 \end{matrix}$

Number of orbit in phosphorous = 3

Electronic configuration of Sulphur (S)

Atomic number of Sulphur = 16



Sulphur

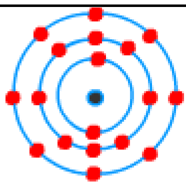
Therefore, number of electrons = 15

Thus, electronic configuration of Sulphur is $\begin{matrix} K & L & M \\ 2 & 8 & 6 \end{matrix}$

Number of orbit in Sulphur = 3

Electronic configuration of Argon (Ar)

Atomic number of argon = 18



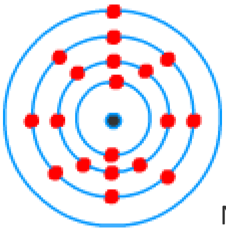
Therefore, number of electrons = 18

Thus, electronic configuration of argon is $\begin{matrix} K & L & M \\ 2 & 8 & 8 \end{matrix}$

Number of orbits in argon = 3

Electronic configuration of Potassium (K)

Atomic number of potassium = 19



Therefore, number of electrons = 19

Since, maximum number of electrons in outermost orbit will not be more than 8, thus the 19th electron of potassium atom will reside in 4th orbit.

Thus, electronic configuration of potassium is $\begin{matrix} K & L & M & N \\ 2 & 8 & 8 & 1 \end{matrix}$

Number of orbits in potassium = 4

Isotopes

Elements having same atomic number but different atomic masses are known as Isotopes.

Example:

Carbon-12, Carbon-13, Carbon-14 are thee isotopes of carbon atom. Here 12, 13 and 14 are the atomic masses of isotopes of carbon respectively. Since, atomic number is the unique property of an atom, thus the atomic number of carbons is 6 even in the case of three types of carbon (isotopes)

126Carbon, 136Carbon, 146Carbon,

Hydrogen-1, Deuterium–2, Tritium-3 are three isotopes of hydrogen.

The isotopes of hydrogen are written as:

11Hydrogen, 21Deuterium, 31Tritium

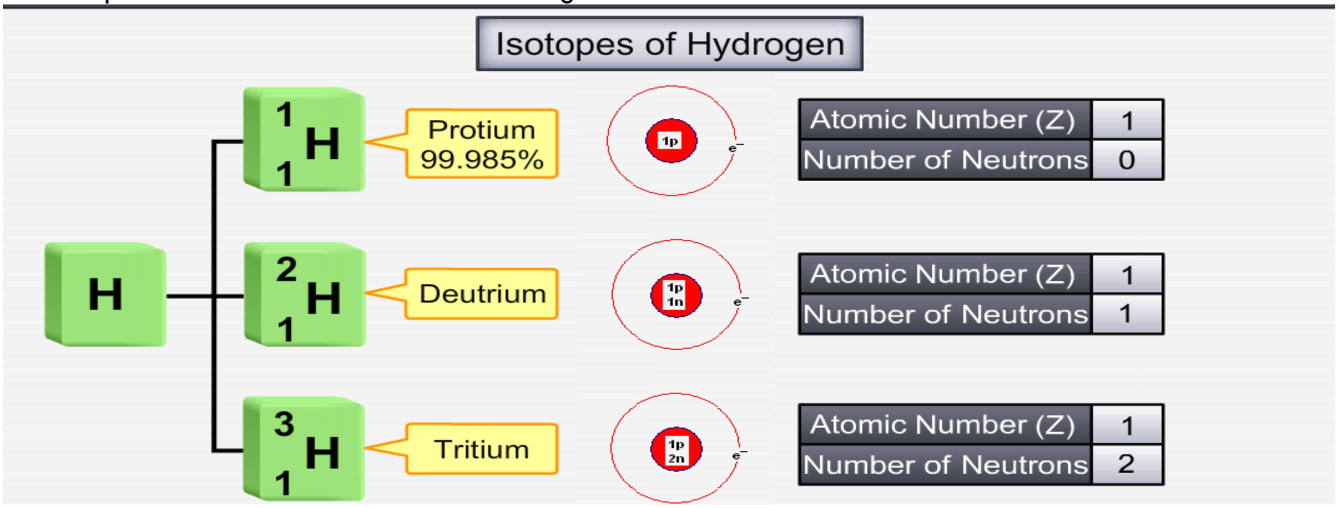
Uses of Isotopes:

Carbon–14 (C-14) is used in carbon dating. This technology is utilized for finding the age of materials found in archaeological excavation. This helps in determining the periods of various events in history. Even fossil's age can be determined by using this technology.

An isotope of uranium is used as fuel in nuclear reactor.

An isotope of cobalt is used in treatment of cancer.

An isotope of iodine is used in treatment of goiter.



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Lesson plan

Atomic theory and structure unit overview

This unit involves the following chemistry concepts.

- Atomic theory
- Organization of the periodic table
- Atomic structure
- Bohr models and Lewis dot diagrams
- Isotopes and ions
- Quantum mechanical model
- Electron configuration and orbital-filling diagrams
- Famous scientists who contributed to atomic theory

Unit rationale

This unit is designed to investigate the development of atomic theory throughout time. A main focus is atomic structure and the quantum model. Notes are provided for electron configuration and orbital filling in case those concepts are a part of your curriculum. If not, adjust the unit as appropriate.

Suggested lesson plan

Lesson 1: The nature of science discussion activity

Objective: The student will explore the philosophical and theoretical nature of science and answer questions based on a conversation in The God Particle *by Leon Lederman.*

Essential knowledge	Plan
• *Constant re-evaluation in the light of new data is essential to ensure scientific knowledge remains current. In this fashion, all forms of scientific knowledge remain flexible and may be revised as new data and new ways of looking at existing data become available.*	1. Distribute *1.1 The God Particle Discussion Questions* before students begin to read. 2. *1.2 The God Particle extract.* Allow 15 minutes for students to read and find answers. 3. Conduct discussion activity using one of the formats from *1.3 God Particle Discussion Instructions for Teachers.* 4. Reflection activity for homework: How do the questions we ask impact on the answers we receive?

Worksheet

WORKSHEET
CLASS IX
TOPIC: STRUCTURE OF ATOM

NAME.______________ ROLL NO ______

1.FILL IN THE BLANKS

_________________ Rutherford's alpha particle scattering experiment led to the discovery of the

Isotopes have same _________________ but different ________________________

Neon and chlorine have atomic numbers 10 and 17 respectively. Their valencies will be _____ and _____

The electronic configuration of Silicon is _______________ and of Sulphur is _________

An atom with 3 protons and four neutrons will have a valency of _____

2.ANSWER THE FOLLOWING

(a)Which particles determine the mass of an atom?

Calculate the number of neutrons present in the nucleus of an element X which is represented as 31 15X.

What is the maximum number of electrons that can be accommodated in the K-Shell of an atom?

Name the scientist who discovered the nucleus.

Name the scientist who discovered that electrons are present in an atom in discrete orbits.

Define isotopes.

Name the particles that determine the mass of an atom.

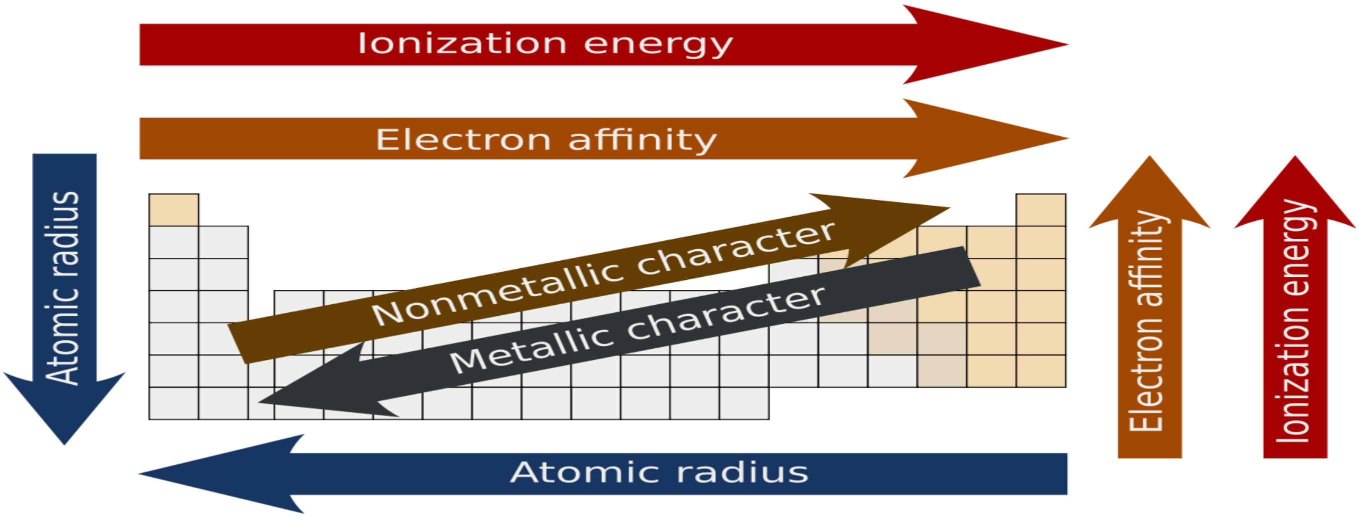
Who proposed the first model of an atom's structure?

What is the number of valence electrons in chloride ion?

The atomic number of fluorine is 9 and its atomic mass is 19. What is the number of neutrons present in its nucleus?

Chapter 03

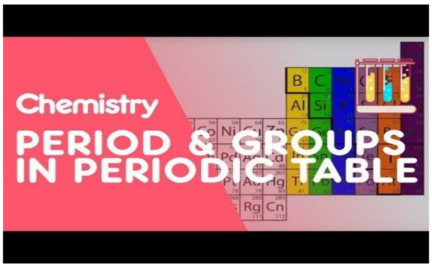
Periodic table and periodicity of properties



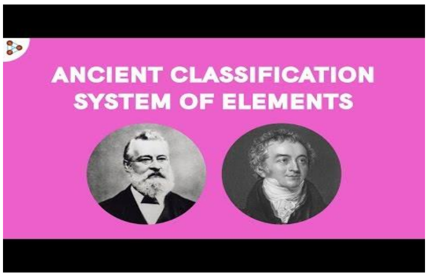
Chapter	Understandings	Skills
Periodic table and periodicity of properties **Introduction of periodic table** Period Groups **Periodicity of properties** Atomic size Ionization energy Electron affinity Electronegativity	Students will be able to: • Distinguish between a period and a group in the periodic table. • State the periodic law. (Remembering) . • Classify the elements (into two categories: groups and periods) • Explain the shape of the periodic table. • Determine the location of families on the Periodic Table. • Identify the relationship between electron configuration and the position of an element on the periodic table. • Describe how electro negativities change within a group and within a period in the periodic table. (Analyzing)	Students will be able to: • Describe how the atomic radii vary within a group and within a period of the periodic table. • Describe how the ionization energies vary within a group and within a period of the periodic table. (Analyzing)

Videos

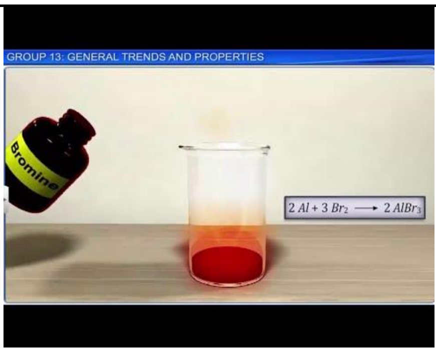
Periodic table



laws of periodic table



Prepared by: Ms. Saima Pervaiz Ahmed under the supervision of Madam Naheed Muneer Siddiqui



S block, P-block ans D-block elements and trends in periodic table

Chapter overview

History of the Periodic Table

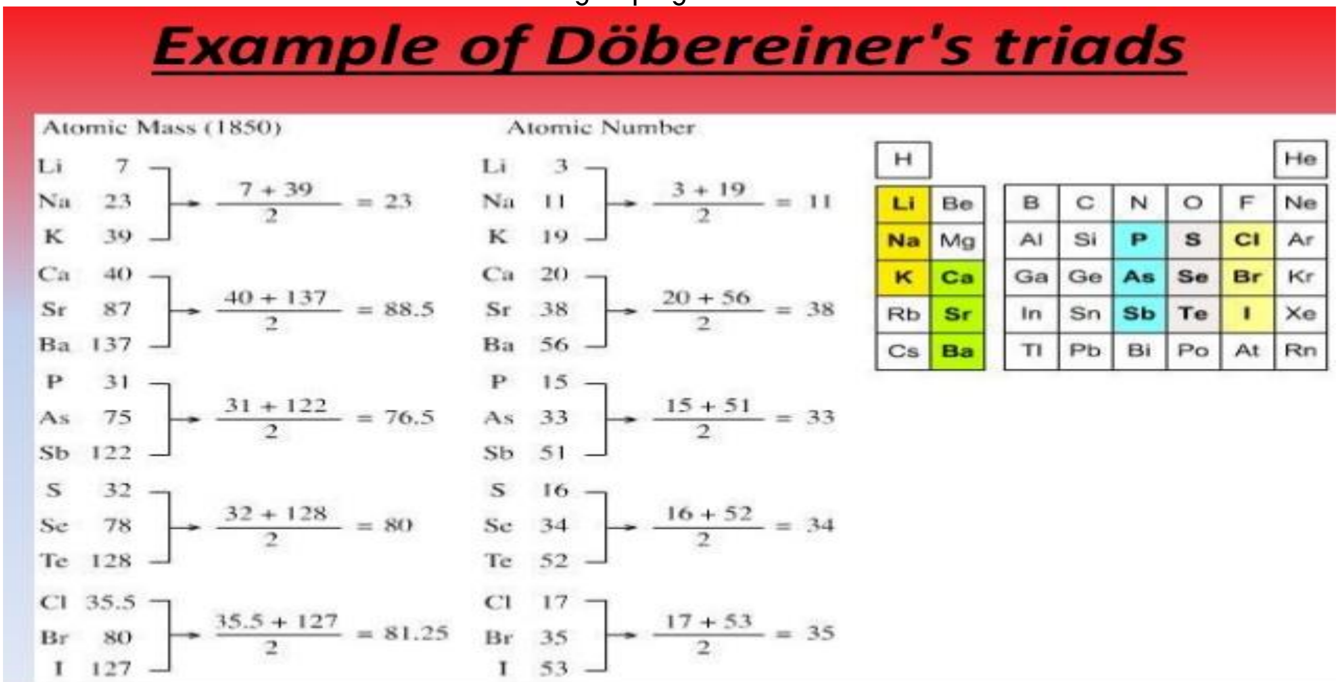
Let us start with the very first theory of periodic table.

1) Doberiener's Triads

The German chemist, Johann Dobereiner in 1800 first observed similarities in the elements on the basis of their properties.

He saw that there are groups consisting of three elements (triads) which have similar chemical and physical properties. Properties of the middle element were also at the halfway of both the elements.

Dobereiner called this grouping method as the law of triads.



For example: Atomic weight of Na = Atomic weight of Li + Atomic weight of K = (7+39)/2 = 23

<u>**The *Limitations of Dobereiner's Triads* are :**</u>

All the elements known at that time couldn't be classified into triads.

Only four triads were mentioned – (Li,Na,K + Ca,Sr,Ba + Cl,Br,I + S,Se,Te).

2) Newland's Octave

In 1865, after the failure of Doberiener's triad the English chemist, John Alexander Newlands gave the law of octaves. According to him, elements can be arranged in ascending order of their atomic weights. He also said that in this arrangement every eighth element of a row had similar properties to that of the first element of the same row, depicting the octaves of music. This law was also dismissed as it was only true for elements up to calcium.

Newlands' Arranged Elements in Octaves:

H	F	Cl	Co/Ni	Br	Pd	I	Pt/Ir
Li	Na	K	Cu	Rb	Ag	Cs	Tl
G	Mg	Ca	Zn	Sr	Cd	Ba/V	Pb
Bo	Al	Cr	Y	Ce/La	U	Ta	Th
C	Si	Ti	In	Zn	Sn	W	Hg
N	P	Mn	As	Di/Mo	Sb	Nb	Bi
O	S	Fe	Se	Ro/Ru	Te	Au	Os

Limitations of Newland's octaves are:

It was only up till calcium that the classification of elements is done via Newland's Octaves.
The discovery of noble gases added to the limitations of this method since they couldn't be included in this arrangement without disturbing it completely.

3) Mendeleev Periodic Table

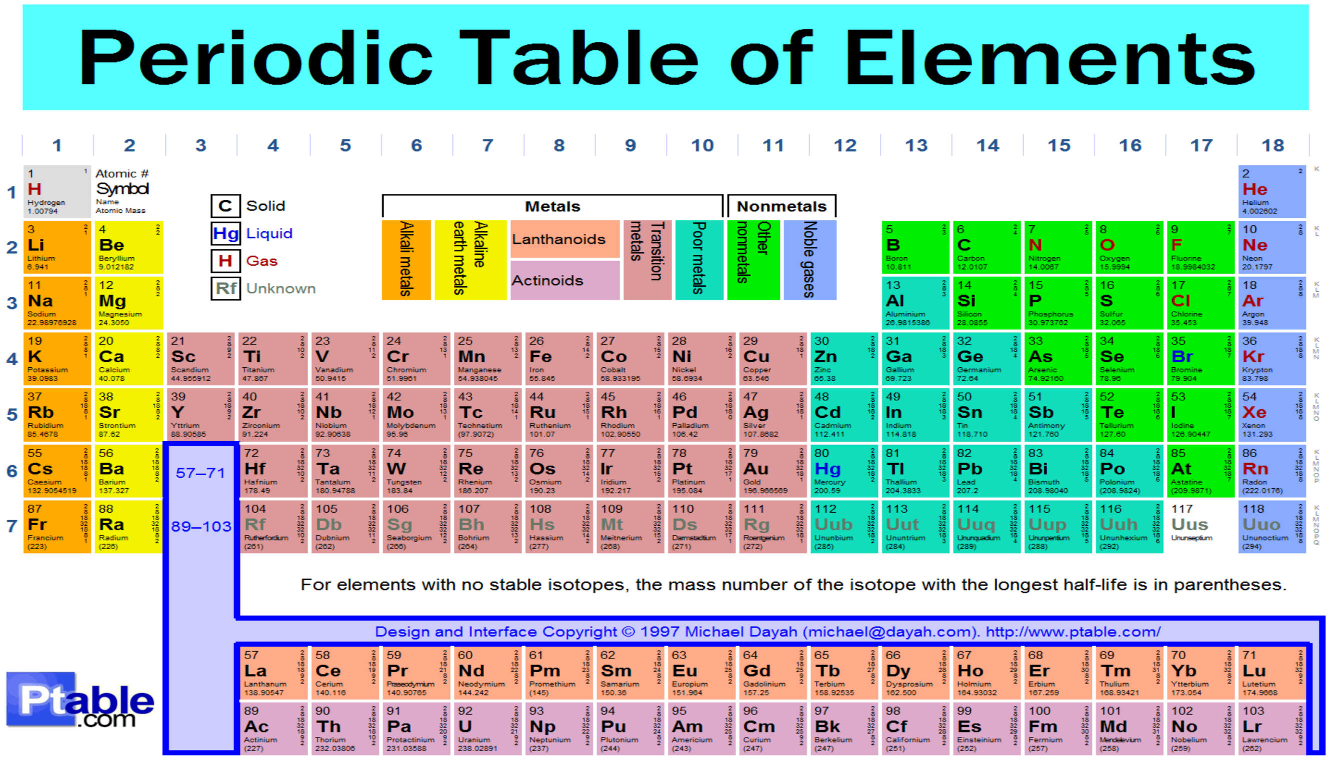
The real development in the periodic table took place after the development of Mendeleev periodic table.
He gave a law which states that "The properties of an element are the periodic function of their atomic masses".
He arranged elements in periods (horizontal rows) and groups (vertical columns) in the increasing order of atomic weights.
The vertical column consists of elements that have similar properties.
Elements are arranged from left to right and top to bottom in the order of their increasing atomic numbers.
Thus;
Elements in the same group will have the same valence electron configuration and hence similar chemical properties.
Whereas, elements in the same period will have an increasing order of valence electrons, therefore, as the energy level of the atom increases, the number of energy sub-levels per energy level increases.
The first 94 elements of the periodic table are naturally occurring, while the rest from 95 to 118 have only been synthesized in laboratories or nuclear reactors.

Modern Periodic Law

The modern periodic law is based on the increasing order of atomic numbers.

A period is a horizontal row of the periodic table.

- There are seven periods in the periodic table, with each one beginning at the far left.
- A new period begins when a new principal energy level begins filling with electrons.
- Period 1 has only two elements (hydrogen and helium)
- While periods 2 and 3 have 8 elements.
- Periods 4 and 5 have 18 elements.
- Periods 6 and 7 have 32 elements because the two bottom rows that are separated from the rest of the table belong to those periods.
- They are pulled out in order to make the table itself fit more easily onto a single page.



A group is a vertical column of the periodic table, based on the organization of the outer shell electrons.

There are a total of 18 groups.

There are two different numbering systems that are commonly used to designate groups and you should be familiar with both.

- Group 1: the alkali metals (lithium family) *not including hydrogen
- Group 2: the alkaline earth metals (beryllium family)
- Groups 3-12: the transition metals
- Group 13: the trials (boron family)
- Group 14: the tetrels (carbon family)
- Group 15: the pnictogens (nitrogen family)
- Group 16: the chalcogens (oxygen family)
- Group 17: the halogens (fluorine family)
- Group 18: the noble gases (helium/neon family

What Is an Element Block?

An element block is a set of elements located in adjacent element groups. Charles Janet first applied the term (in French). The block names (s, p, d, f) originated from descriptions of spectroscopic lines of atomic orbitals.

S-block: The first two groups of the periodic table, the s-block metals:

- Either alkali metals or alkaline earth metals.
- Soft and have low melting points.
- Electropositive and chemically active.

P-block: P-block elements include the last six element groups of the periodic table, excluding helium.

- Include carbon, nitrogen, oxygen, sulfur, halogens, and many other common elements.
- Interact with other chemicals by losing, gaining, or sharing the valence electrons.
- Mostly form covalent compounds (though the halogens form ionic compounds withs-block metals).

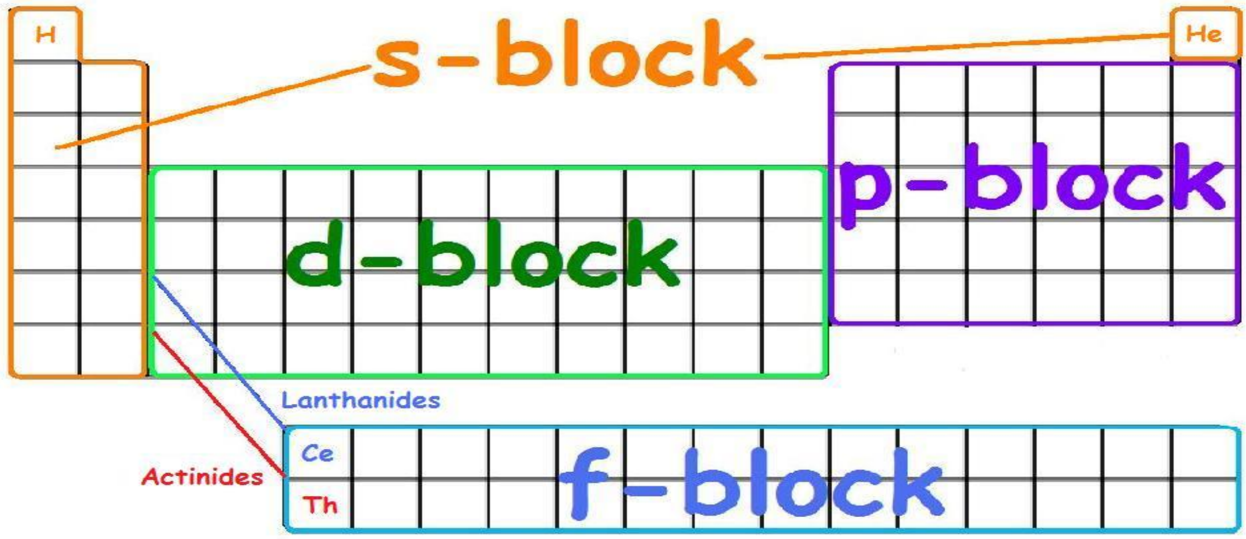
D-block: D-block elements are transition metals of element groups 3-12. D-Block elements: Have valence electrons in their two outermost and shells.

- That is somewhere between that of highly reactive electropositive alkali metals and the covalent compound forming elements (which is why they are called "transition elements").
- Have high melting and boiling points.
- Typically form colored salts.
- Are generally good catalysts.

F-block: Inner transition elements, usually the lanthanide and actinide series, including lanthanum and actinium.

- High melting points.
- Variable oxidations states.
- The ability to form colored salts.

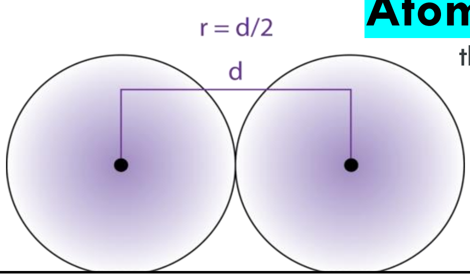
G-block (proposed): G-block would be expected to include elements with atomic numbers higher than 118.



Periodicity of properties

Classification of elements in periodic table

All elements can be classified in one of three classes: metals, metalloids, or nonmetals. Elements in each class share certain basic properties. For example, elements in the metals class can conduct electricity, whereas elements in the nonmetals class generally cannot. Elements in the metalloids class fall in between the metals and nonmetals in their properties.

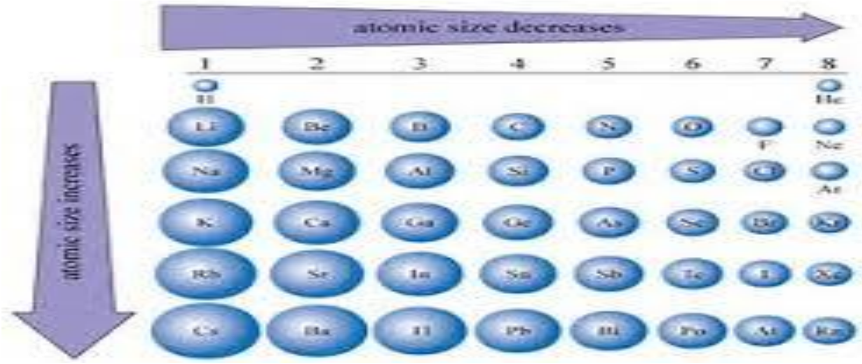


Atomic size The size of atoms is important when trying to explain the behavior of atoms or compounds. One of the ways we can express the size of atoms is with the **atomic radius.**

Atomic radius is determined as the distance between the nuclei of two identical atoms bonded together.

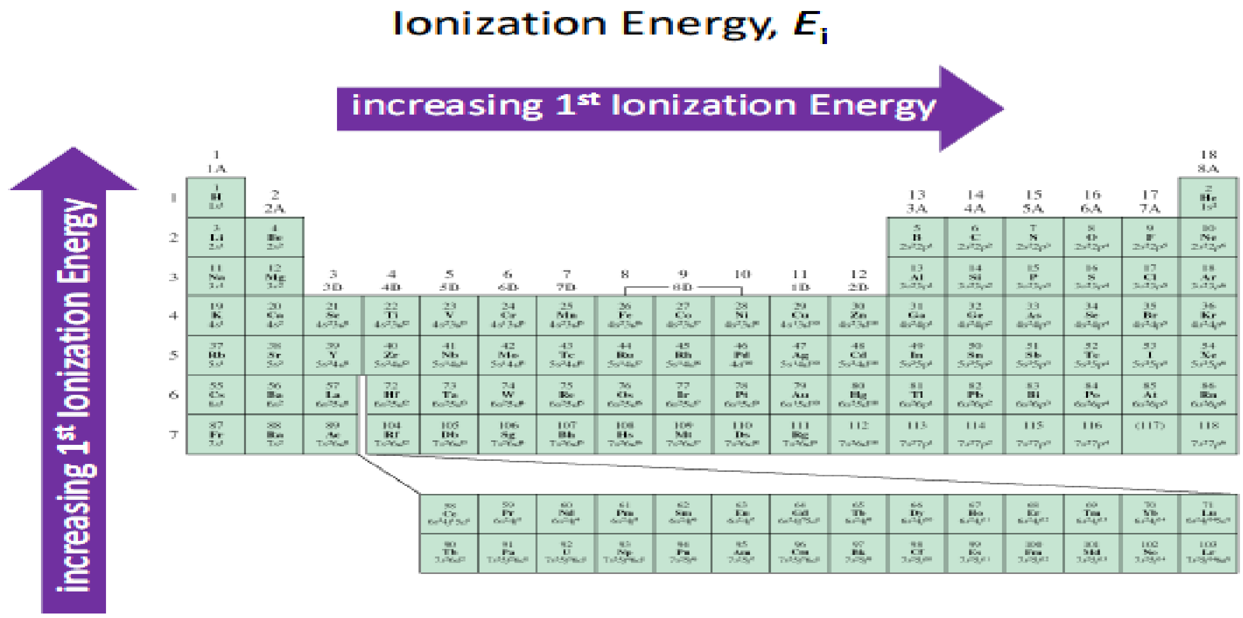
The atomic radius of atoms generally decreases from left to right across a period.

The atomic radius of atoms generally increases from top to bottom within a group.



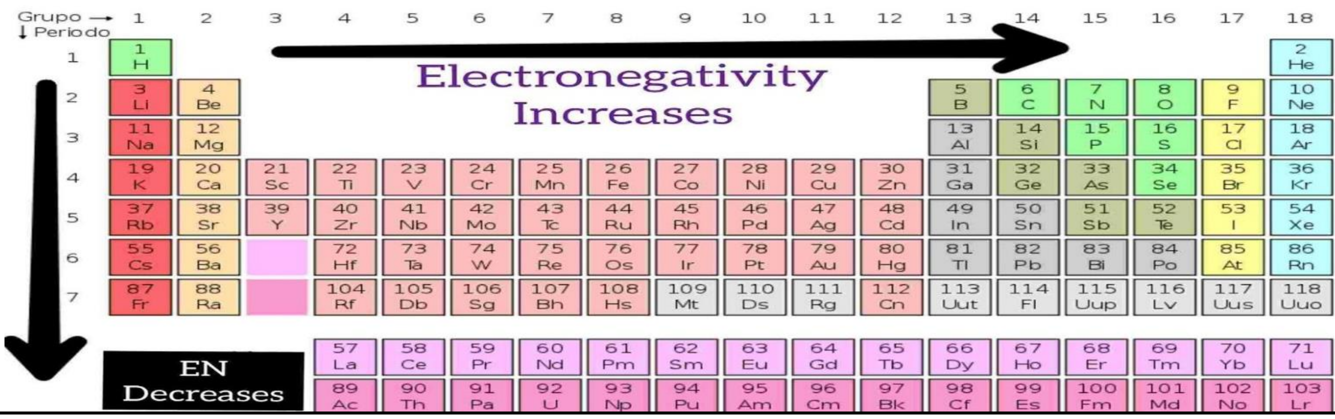
Ionization Energy Trend in the Periodic Table

Ionization energy is the minimum energy required to remove an electron from an atom or ion in the gas phase.
The most common units of ionization energy are kilojoules per mole (kJ/M) or electron volts (eV).



Electronegativity

Electronegativity is a measure of an atom's ability to attract the shared electrons of a covalent bond to itself. If atoms bonded together have the same electronegativity, the shared electrons will be equally shared. If the electrons of a bond are more
As you can see, electro negativities generally increase from left to right across a period and decrease down a group. Attracted to one of the atoms (because it is more electronegative), the electrons will be unequally shared.



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lesson plan

Learning Objectives

- After this lesson, students will be able to:
- Understand that the periodic table is a way to sort elements
- Understand that elements are placed on the periodic table due to similar properties
- Identify a period and a group on the periodic table
- List at least three metals, nonmetals and metalloids
- List at least two properties of each: metals, nonmetals and metalloids

Length

60-85 minutes depending on age group/prior knowledge

Materials

- Depending on the size of your learning group, enough objects for each individual (or team of 2-3) to have ten objects. Objects can be whatever you have available. Examples include: shells, rocks, feathers, cut up fruit, candy, office supplies (tacks, paperclips, staples), donuts, and jellybeans.
- A large piece of butcher paper for each individual (or team) to sort their objects.
- Photocopied periodic tables (one for each student).
- Highlighters and makers (or students can just use pencils).
- A large periodic table that shows metals, nonmetals and metalloids.

Curriculum Standards

Grade 6-8
CCSS.ELA-LITERACY.RST.6-8.4
Determine the meaning of symbols, key terms, and other domain-specific words and phrases as they are used in a specific scientific or technical context relevant to grades 6-8 texts and topics.
CCSS.ELA-LITERACY.RST.6-8.7
Integrate quantitative or technical information expressed in words in a text with a version of that information expressed visually (e.g., in a flowchart, diagram, model, graph, or table).

Worksheet

Periodic Table Facts Worksheet

Name:____ Period _____
Use the following website for help: http://www.chem4kids.com/files/elem_pertable.html

1. Horizontal rows are called __________________.
2. Elements in the same __________________ (row) are filling up the same energy level as you move from left to right.
3. Vertical columns are called __________________.
4. Chemical __________________ have similar properties.
5. __________________ electrons are the electrons in the outermost shell of the atom.
6. All families have the same number of __________________ electrons.
7. The __________________ __________________ have full outer electron shells which makes them inert or non-reactive.
8. Noble gases are the elements in __________________ column (includes helium, neon, __________________, and __________________)
9. Noble gases are elements that have a __________________ valence electron shell.
10. Noble gases are also called __________________ gases because they don't react with other elements.
11. The __________________ family are the most active metals.
12. Alkali metals are the elements in the __________________ column.
13. __________________ is NOT an alkali metal because it behaves unlike any other element and does not belong to a family. It is our only orphan!
14. The __________________ family are the most active NON-metals.
15. Halogens are elements in the __________________ column (includes fluorine, chlorine, __________________, and __________________.

Prepared by: Ms. Saima Pervaiz Ahmed under the supervision of Madam Naheed Muneer Siddiqui

Periodic Table worksheet Name _________________

1. Define a family. ____
2. What is a period? ____
3. What is the symbol for the following elements.

   a. Magnesium _____________ b. Potassium ______________

   c. Iron _______________ d. Copper ____________

4. What are the names of the following elements.

   a. C __________________ b. Cl _________________

   c. Au _________________ d. Sr _________________

5. What period are the following elements in?

   a. He _______________ b. Ge _________________

   c. Rb _______________ d. I __________________

6. What group are the following elements?

   a. Sulfur _______________ b. Ca _________________

   c. Iodine _______________ d. Fe _________________

7. Give me an atom with the following characteristics.

   a. Halogen _________________ b. Nonmetal ________________

   c. Alkali metal ______________ d. metalloid ________________

   e. Lanthanide series __________ f. Alkaline Earth metal ________________

   g. Transition metal ___________ h. Nobel gas ________________

8. Write the electron configuration for

   a. Li ____

   b. Na ____

   c. K ____

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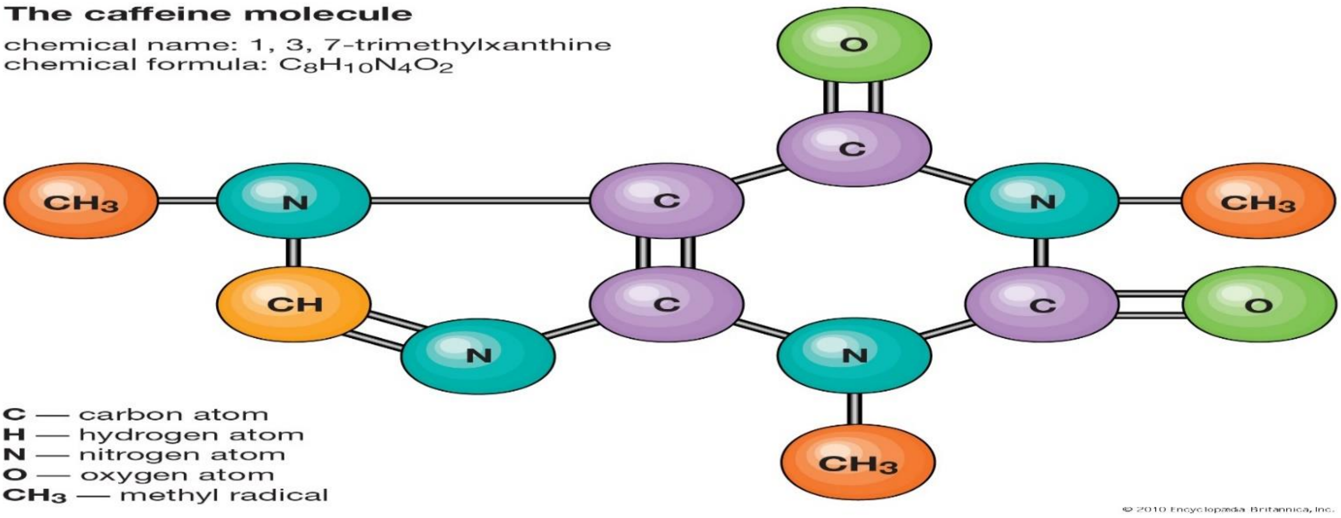
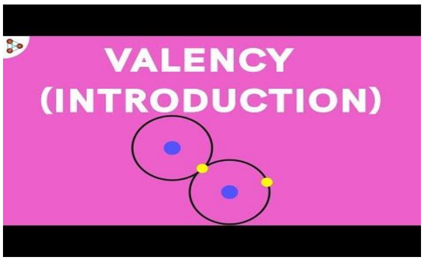
Chapter 04

Structure of molecule



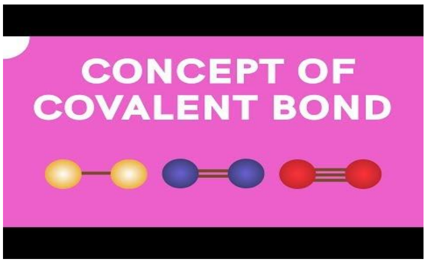
Chapter	Understandings
Structure of molecules **Chemical bonds** **Types of bond** • Ionic bond • Covalent bond • Dative bond • Polar and non polar bond • Metallic bond **Intermolecular forces** • Dipole dipole force • Hydrogen bonding • London force **Nature of bonding and properties** • Ionic compound • Polar and non-polar compound • Metals	Students will be able to: • Students will be able to: • Find the number of valence electrons in an atom using the Periodic Table. • Describe the importance of noble gas electronic configurations. • State the octet and duplet rules. • Explain how elements attain stability. • Describe the ways in which bonds may be formed. • State the importance of noble gas electronic configurations in the formation of ion. • Describe the formation of cations from an atom of a metallic element. • Describe the formation of anions from an atom of a non-metallic element. • Describe the characteristics of an ionic bond. (Understanding). Recognize a compound as having ionic bonds. • Identity characteristics of ionic compounds. • Describe the formation of a covalent bond between two nonmetallic elements. • Describe with examples single, double and triple covalent bonds.

Videos

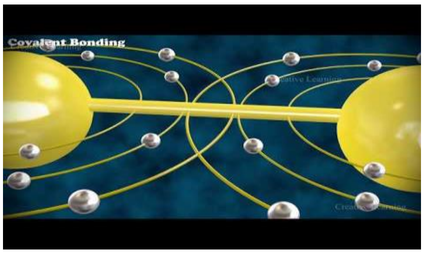


Valancy and bond formation

ionic bond



Covalend bond and its types



Chemical bonding

Metallic bond

Chapter overview

Atom and Molecule

The tiny particle of a chemical element, which may or may not exist independently is called an **atom**.

Molecules refer to the set of the atoms held together by the bond, indicating the smallest unit of a compound.

Bond: A link or force between neighboring atoms in a molecule or compound.

Ions: ions are charge particles. Some atoms become more stable by gaining or losing an entire electron (or several electrons). When they do so, atoms form **ions**, or charged particles.

intermolecular forces: Refers to interactions between two or more molecules

Why atom form chemical bond?

- A ***chemical bond*** is a lasting attraction between atoms, ions or molecules that enables the formation of ***chemical*** compounds.
- Chemical bonds are forces that hold atoms together to make compounds or molecules.
- Chemical bonds include covalent, polar covalent, and ionic bonds.

Types of Chemical Bonds

These types of chemical bonds include:

- Ionic Bonds
- Covalent Bonds
- Hydrogen Bonds
- Polar Bonds

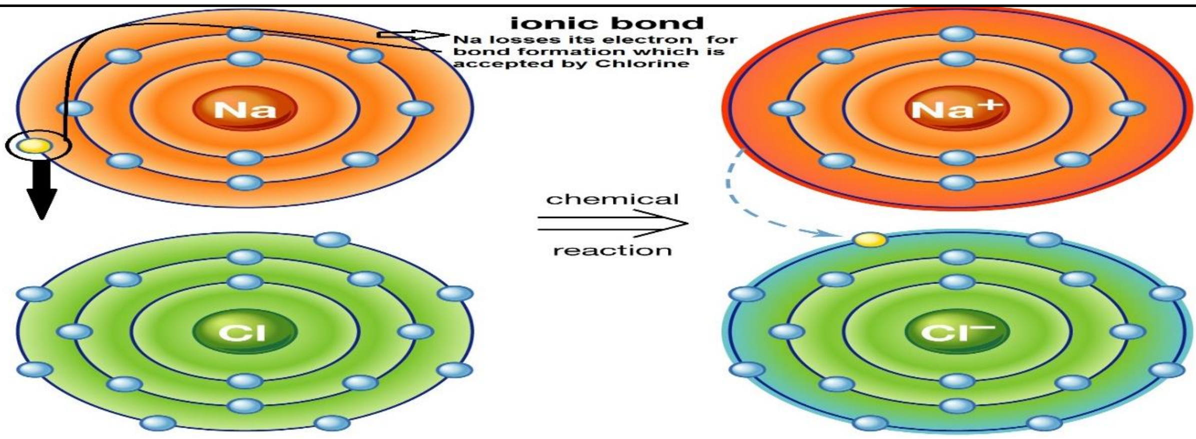
Ionic Bonding

Ionic bonding is a type of chemical bonding which involves a transfer of electrons from one atom or molecule to another. Here, an atom loses an electron which is in turn gained by another atom.
When such an electron transfer takes place, one of the atoms develops a negative charge and is now called the **anion.**

The other atom develops a positive charge and is called the **cation.**

Making an ionic bond

Ionic bonds are bonds formed between ions with opposite charges. For instance, positively charged sodium ions and negatively charged chloride ions attract each other to make sodium chloride, or table salt.
Ionic bonds are formed through the exchange of valence electrons between atoms, typically a metal and a nonmetal.
The loss or gain of valence electrons allows ions to obey the octet rule and become more stable.



Covalent Bonding

A covalent bond indicates the sharing of electrons between atoms. The pair of electrons which are shared by the two atoms now extend around the nuclei of atoms, leading to the creation of a molecule.

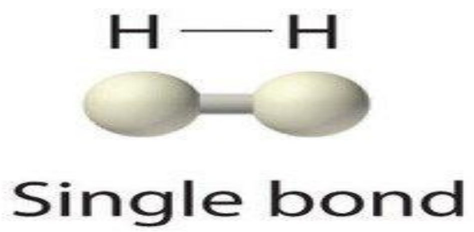
Types of Covalent Bonds

Depending upon the number of shared electron pairs, the covalent bond can be classified into:

- Single Covalent Bond
- Double Covalent Bond
- Triple Covalent Bond

Single Covalent Bonds

A single bond is formed when only one pair of the electron is shared between the two participating atoms. It is represented by one dash (-). Although this form of covalent bond is weaker than a double and triple bond, it is the most stable..

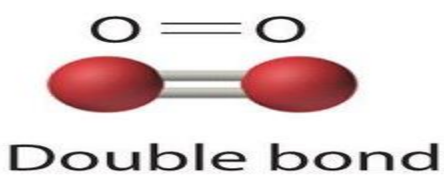


For Example, HCL molecule has one Hydrogen atom with one valence electron and one Chlorine atom with seven valence electrons. In this case, a single bond is formed between hydrogen and chlorine by sharing one electron.

Double Covalent Bonds

A double bond is formed when two pairs of electrons are shared between the two participating atoms. It is represented by two dashes (=). Double covalent bonds are much stronger than a single bond, but they are less stable.

Example: Carbon dioxide molecule has one carbon atom with six valence electrons and two oxygen atom with four valence electrons.

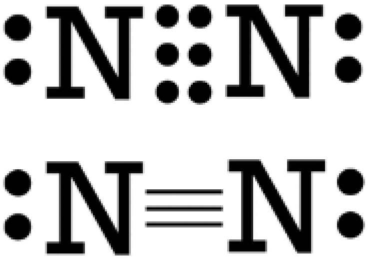


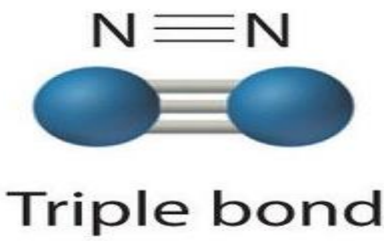
Triple Covalent Bonds

A triple bond is formed when three pairs of electrons are shared between the two participating atoms. Triple covalent bonds are represented by three dashes (≡) and are the least stable types of covalent bonds.

For Example:

In the formation of a nitrogen molecule, each nitrogen atoms having five valence electrons provides three electrons to form three electron pairs for sharing. Thus, a triple bond is formed between the two nitrogen atoms.

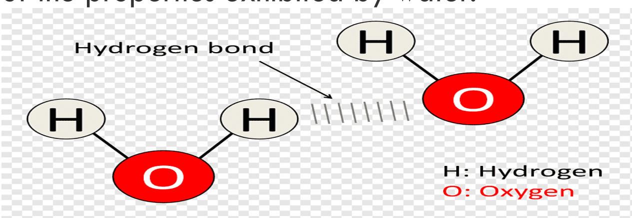




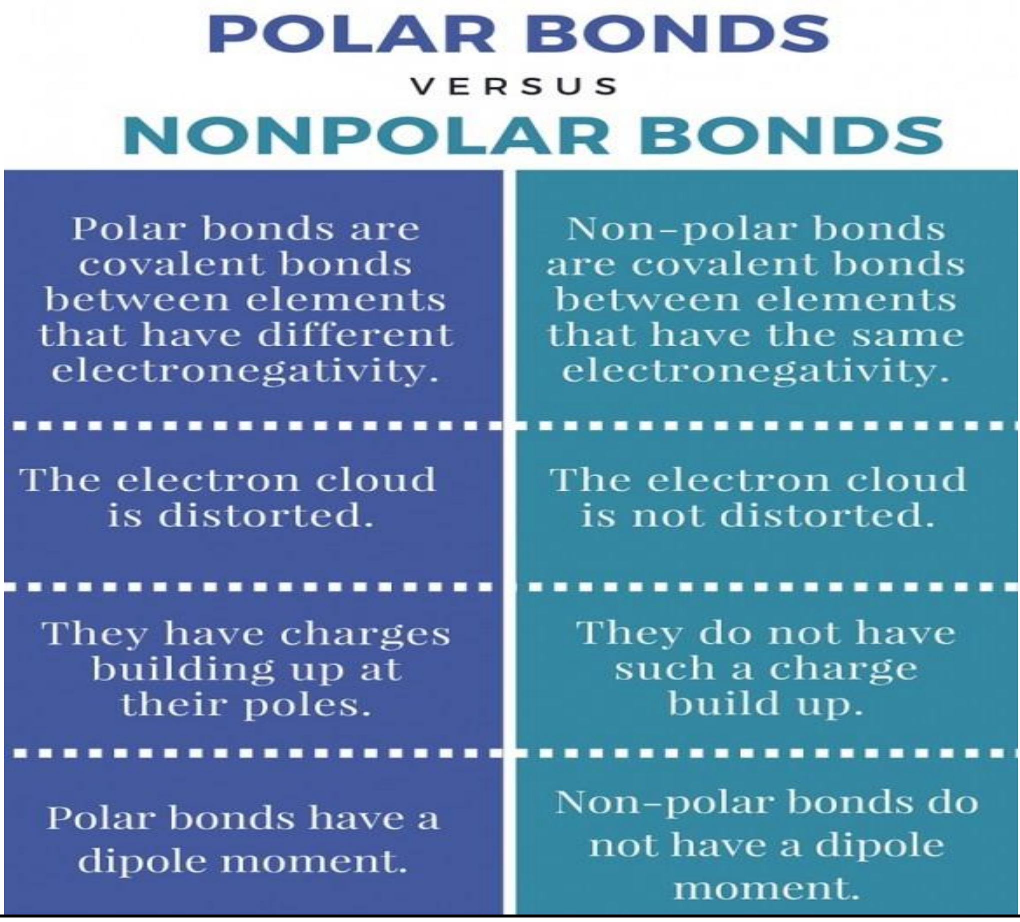
Hydrogen Bonding

Compared to ionic and covalent bonding, Hydrogen bonding is a weaker form of chemical bonding. It is a type of polar covalent bonding between oxygen and hydrogen wherein the hydrogen develops a partial positive charge. This implies that the electrons are pulled closer to the more electronegative oxygen atom.

This creates a tendency for the hydrogen to be attracted towards the negative charges of any neighboring atom. This type of chemical bonding is called a hydrogen bond and is responsible for many of the properties exhibited by water.



Polar and non-polar bond



Dative Bond

A dative bond is a covalent bond between two atoms where one of the atoms provides both electrons that form the bond. A dative bond is also known as a dipolar bond or coordinate bond.

Metallic Bond

A metallic bond is a type of chemical bond formed between positively charged atoms in which the free electrons are shared among a lattice of cations. In contrast, covalent and ionic bonds form between two discrete atoms. Metallic bonding is the main type of chemical bond that forms between metal atoms

Inter molecular forces

Dipole-Dipole Attractions

Dipole–dipole interactions are a type of intermolecular attraction—attractions between two molecules. Dipole-dipole interactions are electrostatic interactions between the permanent dipoles of different molecules. These interactions align the molecules to increase the attraction.

Dispersion Forces or London Forces

It operates for a short distance and it is the weakest force. This kind of force arises due to the movement of electrons thus creating temporary positive and negative charged regions.

Ionic compounds

Ionic compounds are compounds made up of ions. These ions are atoms that gain or lose electrons, giving them a net positive or negative charge. Metals tend to lose electrons, so they become cations and have a net positive charge. Nonmetals tend to gain electrons, forming anions that have a net negative charge.

Reference pages

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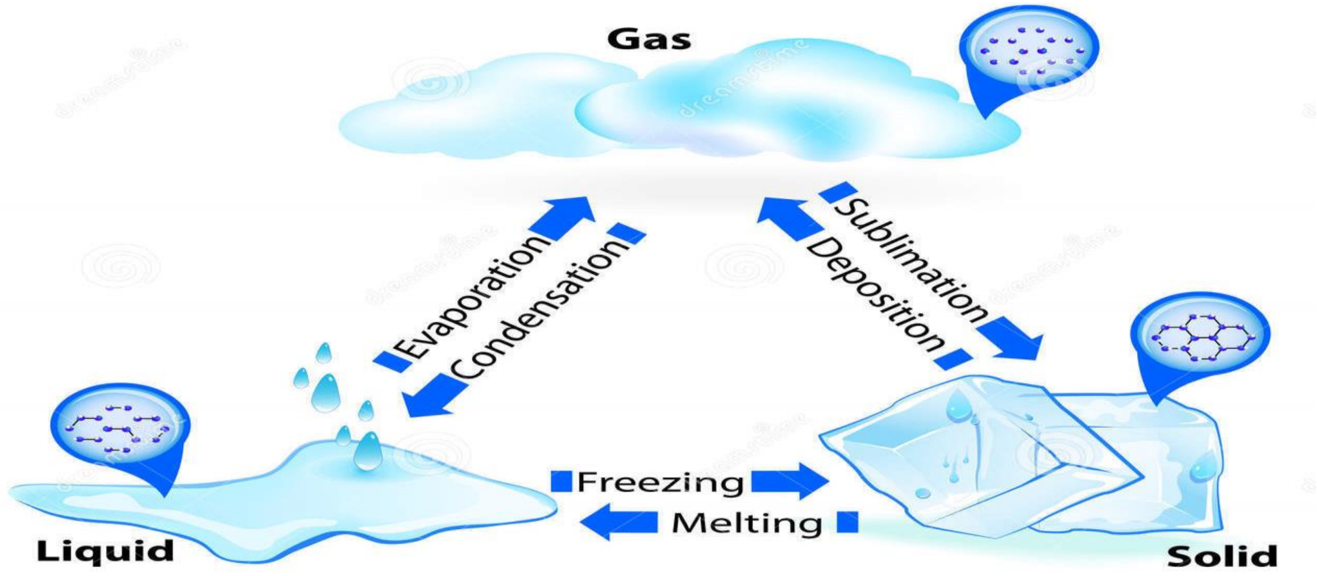
https://courses.lumenlearning.com/trident-boundless-chemistry/chapter/types-of-chemical-bonds/
https://byjus.com/jee/covalent-bond/#:~:text=Types%20of%20Covalent%20Bonds,Triple%20Covalent%20Bond
https://pediaa.com/difference-between-polar-and-nonpolar-bonds/

Lesson plan

https://www.keslerscience.com/chemical-bonding-lesson-plan-a-complete-science-lesson-using-the-5e-method-of-instruction/

Chapter 05

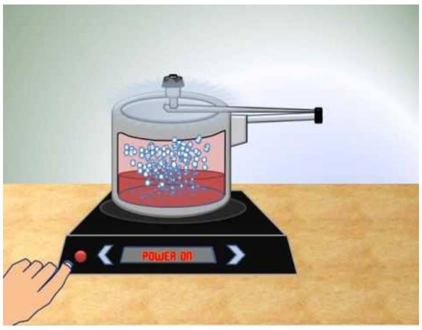
Physical states of matter



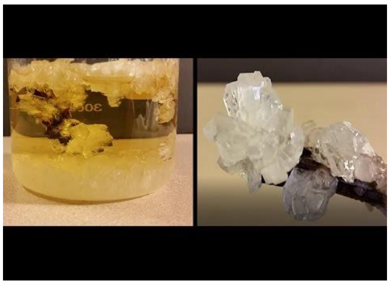
Chapter	Understandings	Skills	Practical
Physical States of Matter • Introduction Gaseous State Typical Properties • (Diffusion, Effusion, Pressure, Compressibility, Mobility, Density) • Laws Related To Gases • Boyle's Law • Charles's Law • Liquid State • Typical Properties (Evaporation, Vapor Pressure, Boiling Point, Freezing Point, Diffusion, Mobility, Density and Factors affecting them.) • Solid State • Typical Properties (Melting Point, Rigidity, Density) • Types of Solids	Students will be able to: • Compare the physical states of matter with regard to intermolecular forces present between them. • Account for pressure-volume changes in a gas using Boyle's Law. (Analyzing) • temperature-volume changes in a gas using Charles's Law. • Explain the properties of gases (diffusion, effusion and pressure). • Summarize the properties of liquids like evaporation, vapor pressure, boiling point. • Describe physical properties of solids	Students will be able to: • Determine melting point of organic solids. • Determine boiling point of organic liquids. • Carry out sublimation. • Carry out distillation process. (Applying)	• Determine the boiling and melting point of benzene or ethyl alcohol. • Calculate the density of solid by using physical balance. • **Grow crystal of sugar (NEW)** • **Preparation of alum crystal. (NEW)**

	(melting and boiling points).		

Videos



States of matter **changes in states of matter** **Gas laws**



Grow crystal of sugar (practical) **Alum crystal formation (Practical)**

Chapter overview

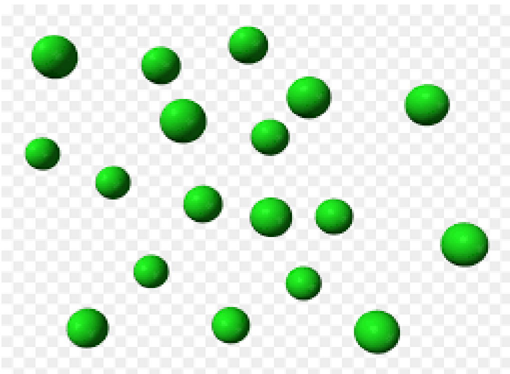
What is Matter?

As discovered by scientists, the matter is made up of very tiny particles and these particles are so small that we cannot see them with naked eyes.
the basic three states of matter:

- **Solid**
- **Liquid**
- **Gas**

The three states of matter are the three distinct physical forms that matter can take in most environments: solid, liquid, and gas.

Gas



- In gases, particles are far apart from each other.
- Force of attraction between the particles is negligible and they can move freely.
- Gases have neither a fixed volume nor a fixed shape.
- The gaseous state has the highest compressibility as compared to solids and liquids.
- The rate is diffusion is higher than solids and liquids.
- The kinetic energy of particles is higher than in solids and liquids.
- An example of gases: air, helium, nitrogen, oxygen, carbon dioxide, etc.

Compressibility and Expandability

The low density of gases makes them compressible since their molecules can be positioned far apart from one another. This allows them to move freely to fit into the gaps of space between them.

Diffusivity

Given the large amounts of space between gas molecules, two or more gases can mix quickly and easily with one another to form a homogeneous mixture. This process is called diffusion.

Pressure

Gas molecules are in constant motion. They exert pressure, or force per unit area, on the interior surface of their container. The pressure varies according to the amount of gas confined to a given container's volume, the temperature and the pressure

Low Density

Gases contain scattered molecules that are dispersed across a given volume and are therefore less dense than in their solid or liquid states.

Boyle's Law

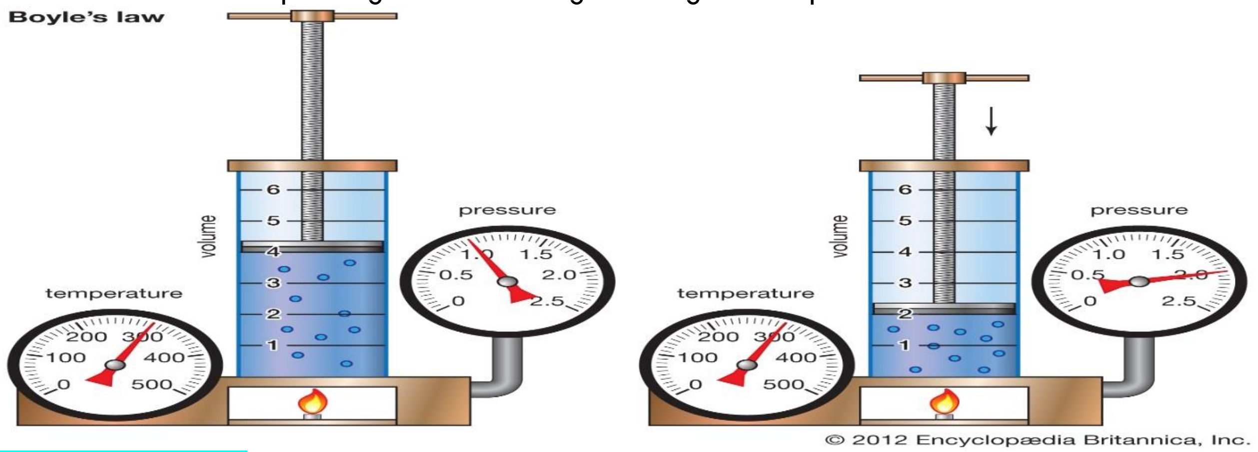
In 1662, Robert Boyle discovered the correlation between *Pressure (P)* and *Volume*

Pressure is inversely proportional to the volume

(V) (assuming *Temperature(T)* and *Amount of Gas(n)* remain constant):

P∝1V→PV=x(Gas Laws.2)

Where x is a constant depending on amount of gas at a given temperature.



Charles' Law

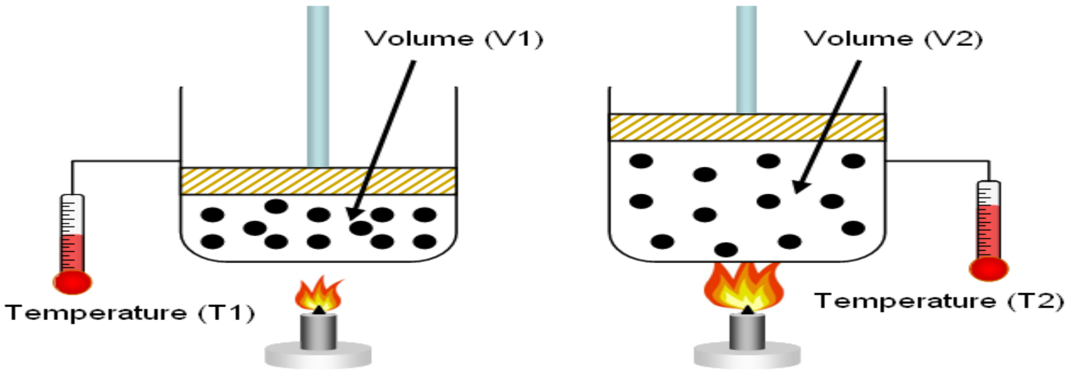
In 1787, French physicists Jacques Charles, discovered the correlation between *Temperature(T)* and *Volume(V)*

Volume is directly proportional to Temperature

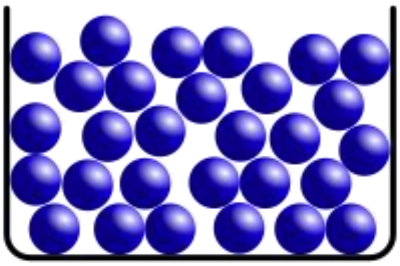
(assuming *Pressure (P)* and *Amount of Gas(n)* remain constant):

V∝T→V=yT(Gas Laws.7)(Gas Laws.7)V∝T→V=yT

Where y is a constant depending on amount of gas and pressure.



Liquid



In a liquid state of matter, particles are less tightly packed as compared to solids.

- Liquids take the shape of the container in which they are kept.
- Liquids are difficult to compress as particles have less space between them to move.
- Liquids have fixed volume but no fixed shape.
- The rate of diffusion in liquids is higher than that of solids.
- Force of attraction between the particles is weak than solids.
- Example of a liquid state of matter: water, milk, blood, coffee, etc.

Changes

A liquid can be converted to a gas through heating at constant pressure to the substance's boiling point or through reduction of pressure at constant temperature. This process of a liquid changing to a gas is called **evaporation**

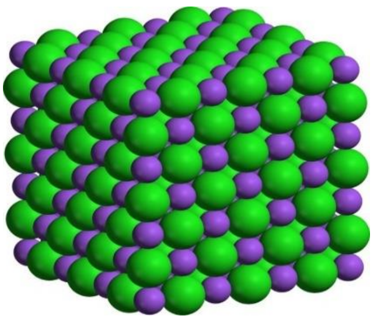
The formation of a gas from a liquid at temperatures below the boiling point is called evaporation.

The vapor pressure for a substance is dependent on the temperature of the substance; as the temperature increases, so does the vapor pressure

The boiling point of a substance is the temperature at which the vapor pressure of a liquid equals the pressure surrounding the liquid and the liquid changes into a vapor.

The freezing point The **temperature** at which a liquid **freezes**: The **freezing point** of water is 32°F, 0°C.

Solid



- In solids, particles are tightly or closely packed.
- The gaps between the particles are tiny and hence it is tough to compress them.
- Solid has a fixed shape and volume.
- Due to its rigid nature, particles in solid can only vibrate about their mean position and cannot move.
- Force of attraction between particles is adamant.
- The rate of diffusion in solids is very low.
- An example of solids: solid ice, sugar, rock, wood, etc

Changes

- A solid can transform into a liquid through melting, and a liquid can transform into a solid through freezing.
- A solid can also change directly into a gas through a process called sublimation.
- Definite shape, definite volume, definite melting point, high density, incompressibility, and low rate of diffusion are the properties of solid

Melting point, temperature at which the **solid** and liquid forms of a pure substance can exist in equilibrium. As heat is applied to a **solid**, its temperature will increase until the **melting point** is reached.

Types of solid

Solids can be classified into two types: crystalline and amorphous.

Crystalline solids are the most common type of solid. They are characterized by a regular crystalline organization of atoms that confer a long-range order.

crystalline: Having a regular three-dimensional molecular lattice structure

Amorphous, or non-crystalline, solids lack this long-range order. Accordingly, they lack the elasticity, distinct melting points, and other properties of crystalline solids.

Amorphous: A solid that lacks the long-range order characteristics of a crystalline solid.

Reference pages

https://byjus.com/chemistry/matter-solid-liquid-gas/
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Lesson plan

https://www.slideshare.net/jamshah/lesson-plan-jamshah

Worksheet

Fill in the blanks:-

1. Matter is made up of small__________.
2. The forces of attraction between the particles are ________ in solids, _______ in liquids and __________ in gases.
3. ___________ is the change of gaseous state directly to solid state without going through liquid state, and vice-versa.
4. Evaporation causes ___________.
5. Latent heat of fusion is the amount of heat energy required to change 1 kg of solid into liquid at its _________.
6. Solid, liquid and gas are called the three ________ of matter.
7. The smell of perfume gradually spreads across a room due to _______.
8. Rapid evaporation depends on the _______ area exposed to atmosphere.
9. As the temperature of a system increases, the pressure of the gases _______.
10. As the volume of a specific amount of gas decreases, it's pressure ________.
11. As the temperature of a gas decreases, I's volume _______.
12. Gas molecules at higher temperatures have more ________ than at cooler temperatures.
13. Usually the total charge of a plasma is ________.
14. The pressure inside of a sealed tube if you raise the temperature go _______
15. Forces of attraction in liquids are ________ than in olid.
16. Liquids that move quickly downhill are described as having _____.

Solution

Fill in the blanks

1. Particles 2. Maximum, intermediate, minimum 3. Sublimation 4. Cooling 5. Melting point
6. States 7. Diffusion 8. Surface 9. Increases 10. Increases 11. Decreases
12. Kinetic energy
13. Zero 14. Up 15. Weaker 16. Low viscosity

True/ False:-

1. Boiling is a bulk phenomenon.
2. Evaporation is a surface phenomenon.
3. The rate of evaporation depends only on the surface area exposed to the atmosphere.

4. Latent heat of vaporization is the heat energy required to change 1 kg. of a liquid to gas at atmospheric pressure at its melting point.
5. Water at room temperature is a liquid.
6. Atoms in a liquid are farther apart than the atoms in a gas.
7. The molecules in a gas are in constant motion.
8. Gases present in air have the same pressure throughout the entire atmosphere.
9. All materials move from solid to liquid to gas as the temperature increases.
10. Because electrons have been stripped away from atoms in plasma, plasmas have a negative charge.
11. It is just as easy to compress a liquid, as it is to compress a gas.
12. Evaporation and boiling are the same processes because molecules move from a liquid to gaseous state.
13. If we pour liquid nitrogen (N2N2) into a glass, it will change its state to a solid.
14. You may find plasma in a star.
15. A system that changes from a solid state to a liquid state gains energy.
16. Plasmas are all made of the same ions. They have different colours due to different amounts of electricity.

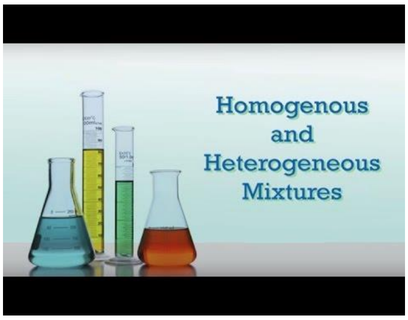
Chapter 06

Solutions

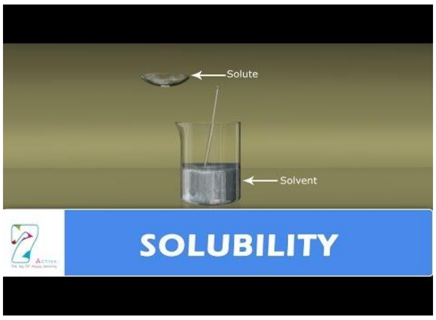
Chapter	Understandings	Skills	Practical
Solution • Solution, Aqueous Solution, • Solute and Solvent • Saturated, Unsaturated, Supersaturated Solutions and Dilution of Solution. • Types of Solution • Solution of Gases (Gases in Gases, Gases In Liquids, Gases in Solids) Solution of Liquids (Liquids in Gases, Liquids in Liquids, Liquids in Solids) Solutions of Solids (Solids in Gases, Solids in Liquids, Solids in Solids) Concentration Units • Percentage • Molarity • Problems Involving the Molarity of a Solution • Solubility • Effect of Temperature on Solubility Comparison of Solutions, Suspension and Colloids	Students will be able to: • Define the terms: solution, aqueous solution, solute and solvent and give an example of each. • Explain the difference between saturated, unsaturated and supersaturated solutions. Explain the formation of solutions • Explain what is meant by the concentration of a solution. • Define Molarity. • Define percentage solution. • Solve problems involving the Molarity of a solution. (Applying) • Describe how to prepare a solution of given Molarity. (Applying)	Students will be able to: • Prepare solutions of different strength. (Applying) • Carry out dilution of solutions. (Applying	• Prepare 100 cm3 of 0.1 M Na2CO3 solution. • Prepare 250 cm3 of 0.1M of oxalic acid solution. • Determine the solubility of NaCl salt.

Videos

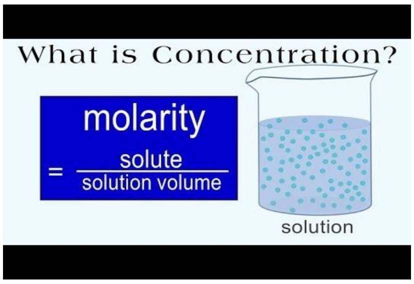
Solution, suspension and colloids



homogenous & heterogeneous Mixtures



Solubility



Molarity

Chapter overview

Define Solution

A solution is defined as

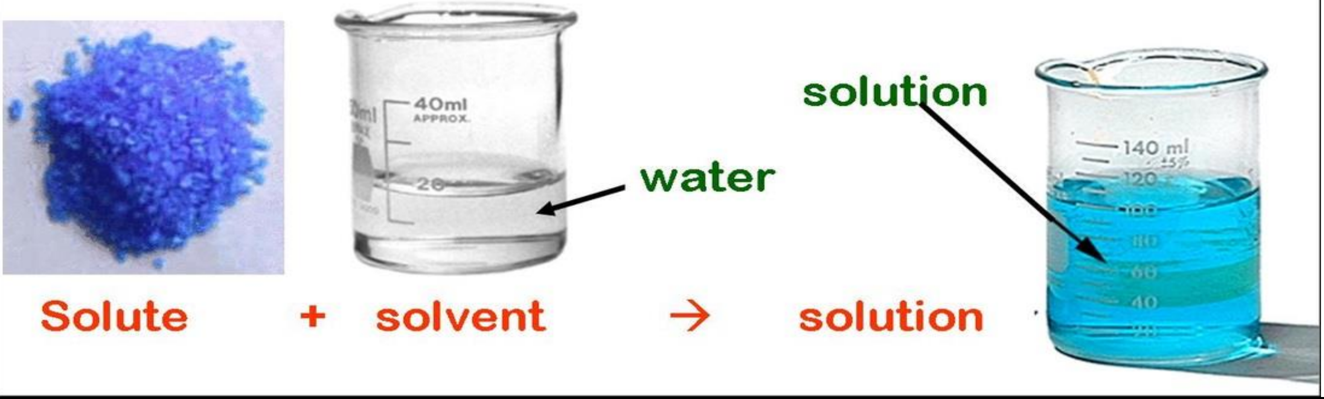
A homogenous mixture which mainly comprises of two components namely solute and solvent. For example, salt and sugar is a good illustration of a solution. A solution can be categorized into several components.

The two main parts of a solution. These are:

Solute: this is the substance that makes up the minority of the solution, or this is the part that is dissolved. In our example of salt water, the solute is the salt.

Solvent: this is the substance that makes up the majority of the solution. This is the part where the solute is dissolved. In our example of salt water, the solvent is water.

A homogeneous mixture is a type of mixture with a uniform composition. This means that the substances cannot be distinguished easily from one another.



Different Types of Solutions

On the Basis of Water as Solvent

Based on the whether the solvent is water or not, solutions are of two types.

Aqueous solutions: These solutions have water as the solvent. Examples of such solutions are sugar in water, carbon dioxide in water, etc.

Non-Aqueous Solutions: These solutions have a solvent that is not water. It could be ether, benzene, petrol, carbon tetrachloride etc. Common examples include sulfur in carbon disulphide, naphthalene in benzene, etc.

On the Basis of the Amount of Solute Added

Based on the amount of solute present in the solution, we can classify them into the following types.

Unsaturated Solution: An unsaturated is one that can dissolve more solute at a definite temperature. It means that we can still add more solute to the solvent.

Saturated Solution: A solution is said to be saturated when we can't add any more solute to the solvent. This means that the solution can't dissolve any more solute at a definite temperature.

Supersaturated Solution: A supersaturated solution is one where the solute is present in an excess amount. This solute is dissolved forcefully by raising the temperature or pressure of the solution. These generally crystal out in the bottom by the method called crystallization.

On the Basis of Amount of Solvent Added

Concentrated Solution: A concentrated solution has large amounts of solute in the given solvent. Examples include Brine solution, Orange juice, and dark colour tea.

Dilute Solution: A dilute solution has a small amount of solute in a large amount of solvent. Examples include Salt solution, light colour tea.

Solid , liquid and gaseous solutions

Solute	Solvent	Solution
Gaseous Solutions		
gas	gas	air (nitrogen, oxygen, argon gases)
liquid	gas	humid air (water vapor in air)
Liquid Solutions		
gas	liquid	carbonated drinks (CO_2 in water)
liquid	liquid	vinegar (CH_3COOH in water)
solid	liquid	salt water (NaCl in water)
Solid Solutions		
liquid	solid	dental amalgam (Hg in Ag)
solid	solid	sterling silver (Cu in Ag), alloys

Concentration Units based on moles

There are a number of ways to express the relative amounts of solute and solvent in a solution. This page describes calculations for four different units used to express concentration:

- Percent Composition (by mass)
- Molarity
- Molality
- Mole Fraction

Percent Composition (by mass)

We can consider percent by mass (or weight percent, as it is sometimes called) in two ways:
The parts of solute per 100 parts of solution.
The fraction of a solute in a solution multiplied by 100.
We need two pieces of information to calculate the percent by mass of a solute in a solution:
The mass of the solute in the solution.
The mass of the solution.
Use the following equation to calculate percent by mass:

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

Molarity

Molarity tells us the number of moles of solute in exactly one liter of a solution. (Note that molarity is spelled with an "r" and is represented by a capital M.)
We need two pieces of information to calculate the molarity of a solute in a solution:
The moles of solute present in the solution.
The volume of solution (in liters) containing the solute.
To calculate molarity we use the equation:

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solution in liters}}$$

Example: What is the molarity when 0.750 mol is dissolved in 2.50 L of solution?
Solution:

$$\text{Molarity} = \frac{0.750 \text{ mol}}{2.50 \text{ L}} = 0.300 \text{ M}$$

Example: 80.0 grams of glucose ($C_6H_{12}O_6$, mol. wt = 180. g/mol) is dissolved in enough water to make 1.00 L of solution. What is its molarity?
Solution:
1) Convert grams to moles:

$$\frac{80.0 \text{ g}}{180.0 \text{ g/mol}} = 0.444 \text{ mol}$$

2) Calculate the molarity:

$$\frac{0.444 \text{ mol}}{1.00 \text{ L}} = 0.444 \text{ M}$$

Molality

Molality, m, tells us the number of moles of solute dissolved in exactly one kilogram of solvent. (Note that molality is spelled with two "l"'s and represented by a lower case m.)
We need two pieces of information to calculate the molality of a solute in a solution:
The moles of solute present in the solution.
The mass of solvent (in kilograms) in the solution.
To calculate molality we use the equation:

$$\text{Molality} = \frac{\text{moles of solute}}{\text{mass of solvent in kilograms}}$$

Example: Suppose you had 2.00 moles of solute dissolved into 1.00 L of solvent. What's the molality?

$$\text{Molality} = \frac{2.00 \text{ mol}}{1.00 \text{ kg}}$$

The answer is 2.00 m.
Notice that no mention of a specific substance is mentioned at all. The molarity would be the same no matter what the substance. It doesn't matter if it is sucrose, sodium chloride or any other substance. One mole of anything contains 6.022×10^{23} units.

Mole Fraction

The mole fraction, *X*, of a component in a solution is the ratio of the number of moles of that component to the total number of moles of all components in the solution.
To calculate mole fraction, we need to know:
The number of moles of each component present in the solution.
The mole fraction of A, X_A, in a solution consisting of A, B, C, ... is calculated using the equation:

Solubility

In chemistry, a **suspension** is a heterogeneous mixture that contains solid particles sufficiently large for sedimentation. The particles may be visible through naked eye, usually must be larger than one micrometer,

Factors Affecting Solubility

Solid Solubility and Temperature

For many solids dissolved in liquid water, the solubility increases with temperature
The solubility of a given solute in a given solvent typically depends on temperature. Many salts show a large increase in solubility with temperature.

Gas Solubility and Temperature

Solubility of a gas in water tends to decrease with increasing temperature, and solubility of a gas in an organic solvent tends to increase with increasing temperature.

Solubility and Pressure

Increasing pressure will increase the solubility of a gas in a solvent.

Henry's law: William Henry, an English chemist, showed that the solubility of a gas increased with increasing pressure. States that the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid.
$P \propto C$ (or) $P = k_H.C$
Where,

- 'P' denotes the partial pressure of the gas in the atmosphere above the liquid.

- 'C' denotes the concentration of the dissolved gas.
- 'k_H' is the Henry's law constant of the gas.

Calculate the solubility of gaseous oxygen in water at a temperature of 293 K when the partial pressure exerted by O_2 is 1 bar. (Given: k_H for O_2 34840 bar.L.mol^{-1})

As per Henry's law, P = k_H*C
Substituting, k_H = 34840 bar.L.mol^{-1} and P = 1 bar, the equation becomes
C = 1/34840 mol.L^{-1} = 2.87*10^{-5} mol/L
Therefore, the solubility of oxygen in water under the given conditions is 2.87*10^{-5} M

Suspension

A suspension is a heterogeneous mixture of two or more substances. In it, the particles are suspended throughout the solution in bulk and can be easily seen by naked eyes.
Some of the common examples of include:

- Mixture of chalk and water
- Muddy water
- Mixture of flour and water
- Mixture of dust particles and air
- Fog
- Milk of magnesia, etc.

Colloids

In chemistry, a colloid is a mixture in which one substance of microscopically dispersed insoluble or soluble particles is suspended throughout another substance.

Reference pages

https://byjus.com/chemistry/types-of-solutions/
https://study.com/academy/lesson/what-is-a-solution-in-science-definition-examples.html#:~:text=A%20solution%20is%20a%20homogeneous,the%20majority%20of%20the%20solution.
https://www.toppr.com/guides/chemistry/solutions/types-of-solutions/
https://www.toppr.com/guides/chemistry/solutions/types-of-solutions/
https://www.chem.purdue.edu/gchelp/howtosolveit/Solutions/concentrations.html
https://byjus.com/chemistry/henrys-law/
https://courses.lumenlearning.com/boundless-chemistry/chapter/factors-affecting-solubility/
https://www.chemteam.info/Solutions/Molality.html#:~:text=The%20molality%20of%20a%20solution,by%20the%20kilograms%20of%20solvent.&text=This%20is%20probably%20easiest%20to,into%20exactly%201.00%20liter%20water.

Lesson plan

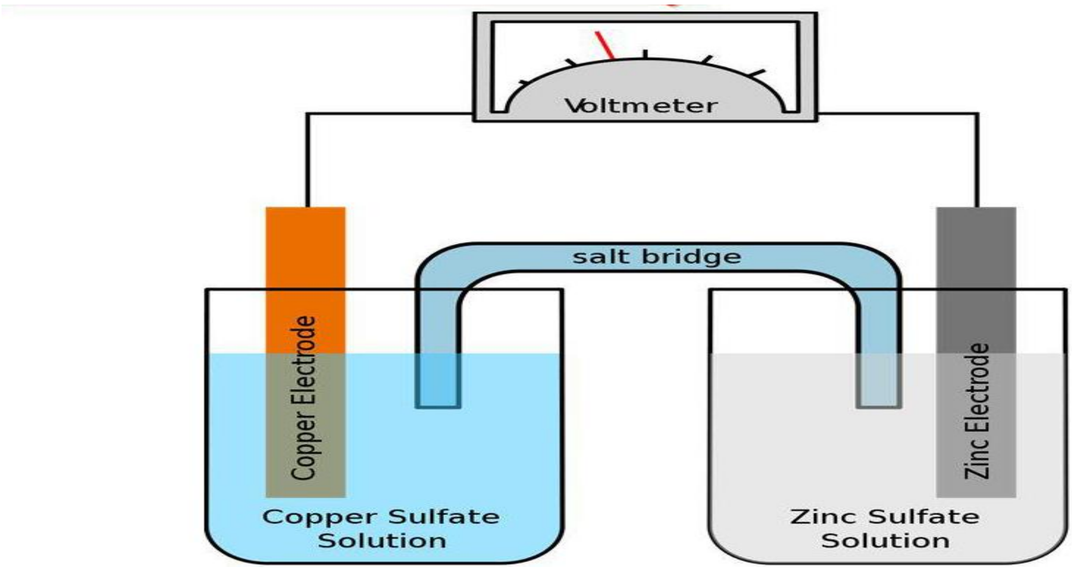
https://aminghori.blogspot.com/2016/12/lesson-plan-of-difference-between.html

Worksheet

1. What are the three different types of mixtures?
2. What is a solution?
3. Classify each of the following as a heterogeneous mixture or a homogeneous mixture.
a. Salad ________________________________
b. Tap water ____________________________
c. Muddy water __________________________
4. What is the different between a solute and solvent?
5. What is considered to be the 'universal' solvent?
6. Not all solutions are solids dissolved in liquids. Provide two examples of other types of solutions.
7. What is the Tyndall Effect? Give a common example of this effect.
8. In what type of mixture is to easiest to separate the component substance? Why?
9. Given an unknown mixture consisting of two substances, explain how a scientist could use lab techniques to determine whether the mixture is a true solution, a colloid, or a suspension.

Chapter 07

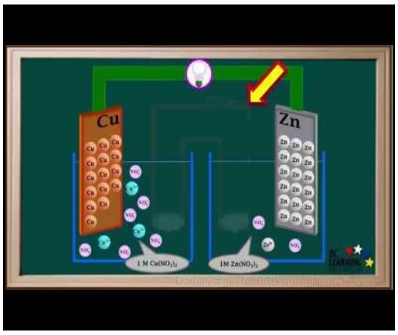
Electrochemistry



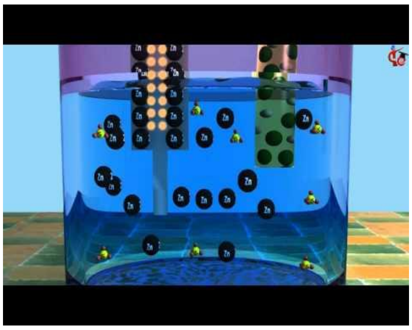
Chapter	Understandings	Skills	Practical
Electrochemistry • Introduction • **Oxidation and Reduction** Oxidizing and Reducing Agents Oxidation - Reduction Reactions Electrochemical Cells • **Concept of Electrolytes** • **Faraday's 1st and 2nd law of electrolysis** • **Electrolytic Cells** • **Galvanic Cells** (Daniel Cell) Electrochemical Industries Manufacture of Sodium Metal from Fused NaCl • **Batteries** • Corrosion and Its Prevention Rusting of Iron • Electroplating of Tin, Zinc, Silver and Chromium on Steel	Students will be able to: • Define oxidation and reduction in terms of loss or gain of oxygen or hydrogen. (Understanding) • Define oxidation and reduction in terms of loss or gain of electrons. (Understanding) • Define oxidizing and reducing agents in a redox reaction. (Understanding) Define oxidation state. (Remembering) • State the common rules used for assigning oxidation numbers to free elements, ions (simple and complex), molecules, atoms.(Remembering) • Describe the nature of electrochemical processes. (Understanding)	Students will be able to: • Determine which solutions conduct electricity given a set of different solutions. (Analyzing) • Perform metal displacement reactions in aqueous medium. (Applying)	• Demonstrate the conductivity of different given solution • Demonstrate the metal displacement reaction in aqueous solution.

	• Identify the direction of movement of cations and anions towards respective electrodes. (Understanding) • List the possible uses of an electrolytic cell. (Understanding) • Sketch a Danniell cell, labeling the cathode, the anode, and the direction of flow of the electrons. (Understanding)		

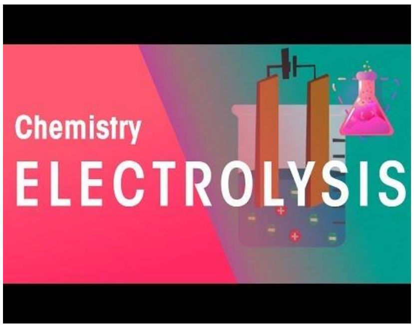
Videos



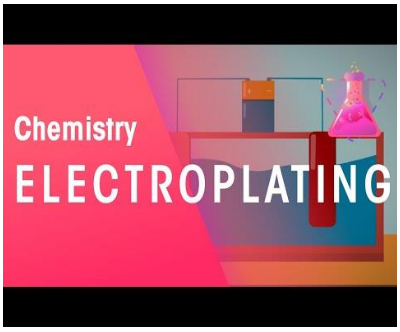
Electrochemical cell



Glavanic cell

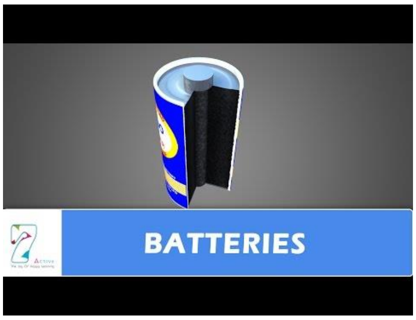


Electrolysis



Electroplating

Redox reaction



Batteries

Chapter overview

OXIDATION

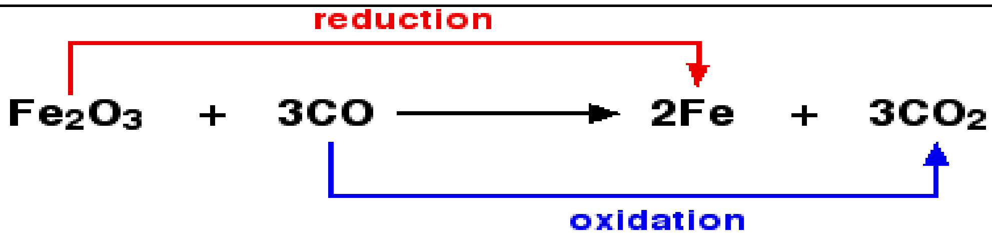
Oxidation is the loss of electrons during a reaction by a molecule, atom or ion. ***Oxidation*** occurs when the ***oxidation*** state of a molecule, atom or ion is increased.

REDUCTION

Reduction is a chemical reaction that involves the gaining of electrons by one of the atoms involved in the reaction between two chemicals.

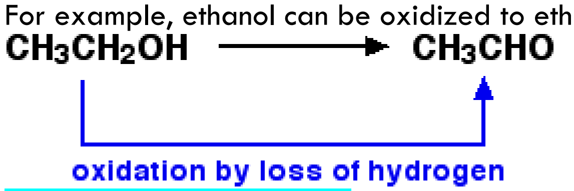
Oxidation and Reduction with respect to Oxygen Transfer

Oxidation is the **gain** of oxygen.
Reduction is the **loss** of oxygen.
For example, in the extraction of iron from its ore:



Oxidation and reduction in terms of hydrogen transfer

Oxidation is the **loss** of hydrogen.
Reduction is the **gain** of hydrogen.
For example, ethanol can be oxidized to ethanol:



Redox Reaction

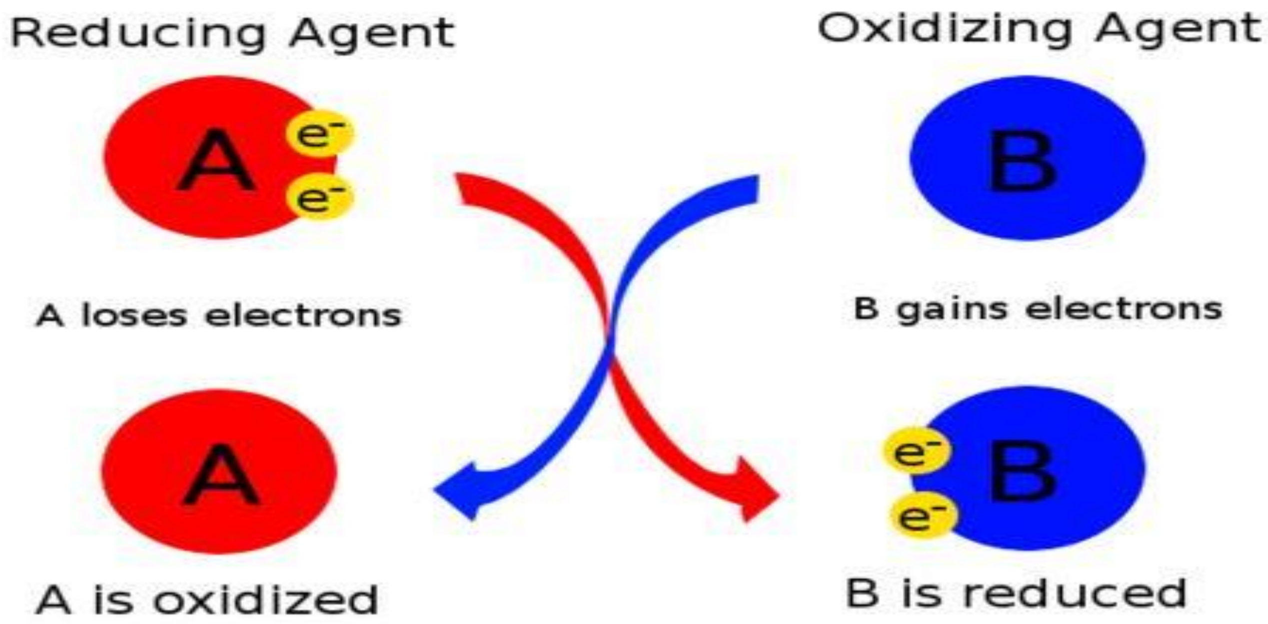
Both reduction and oxidation are occurring simultaneously, this is known as a **redox reaction** An oxidation-reduction (redox) reaction is a type of chemical reaction that involves a transfer of electrons between two species.
An oxidation-reduction reaction is any chemical reaction in which the oxidation number of a molecule, atom, or ion changes by gaining or losing an electron. Redox reactions are common and vital to some of the basic functions of life, including photosynthesis, respiration, combustion, and corrosion or rusting.

Oxidizing Agent

An oxidizing agent is a reactant that removes electrons from other reactants during a redox reaction. The oxidizing agent typically takes these electrons for itself, thus gaining electrons and being reduced. An oxidizing agent is thus an electron acceptor.

Reducing Agent

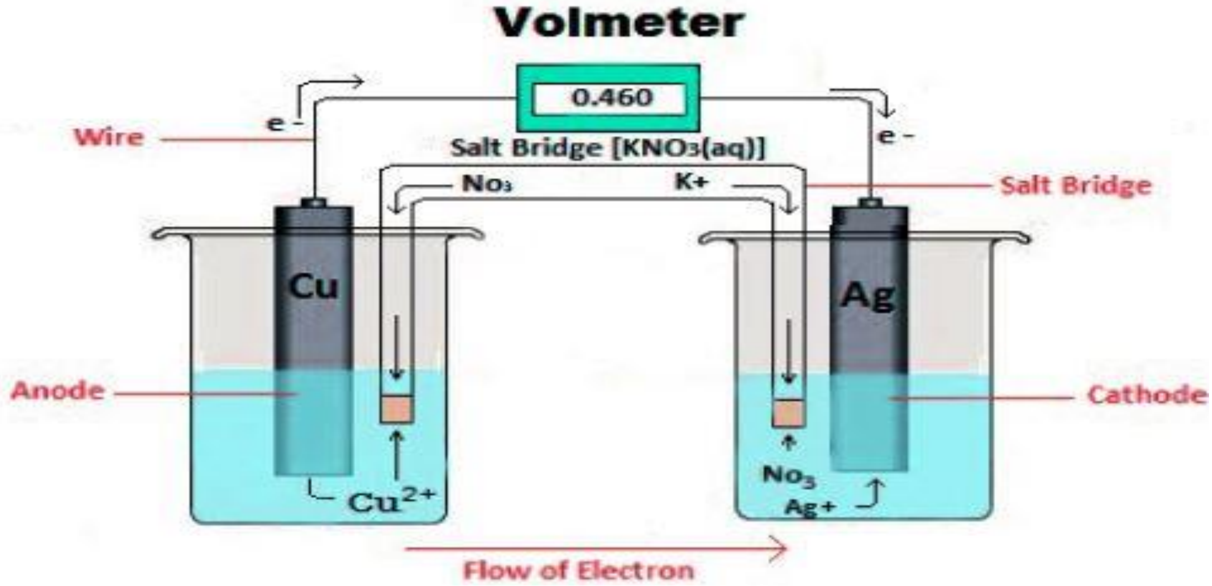
A reducing agent is ***one of the reactants of an oxidation-reduction reaction which reduces the other reactant by giving out electrons to the reactant***. If the reducing agent does not pass electrons to other substance in a reaction, then the reduction process cannot occur



What is an Electrochemical Cell?

An electrochemical cell is a device that can generate electrical energy from the chemical reactions occurring in it, or use the electrical energy supplied to it to facilitate chemical reactions in it.

These devices are capable of converting chemical energy into electrical energy, or vice versa. A common example of an electrochemical cell is a standard 1.5-volt cell which is used to power many electrical appliances such as TV remotes and clocks.



Types of Electrochemical Cells

The two primary types of electrochemical cells are
1. Galvanic cells (also known as Voltaic cells)
2. Electrolytic cells

The key differences between Galvanic cells and electrolytic cells are tabulated below.

Galvanic Cell / Voltaic Cell	Electrolytic Cell
Chemical energy is transformed into electrical energy in these electrochemical cells.	Electrical energy is transformed into chemical energy in these cells.
The redox reactions that take place in these cells are spontaneous in nature.	An input of energy is required for the redox reactions to proceed in these cells, i.e. the reactions are non-spontaneous.
In these electrochemical cells, the anode is negatively charged and the cathode is positively charged.	These cells feature a positively charged anode and a negatively charged cathode.
The electrons originate from the species that undergoes oxidation.	Electrons originate from an external source (such as a battery).

What is Galvanic Cell?

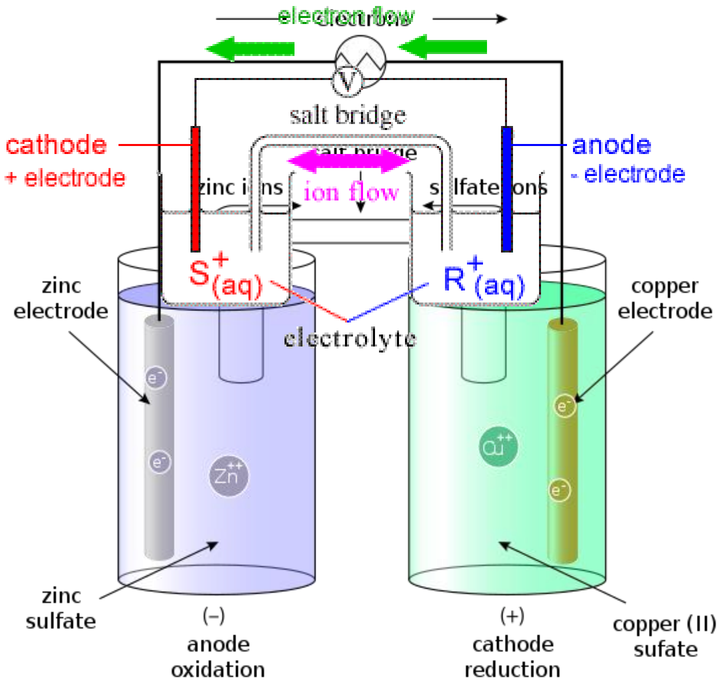
An electrochemical cell that converts the chemical energy of spontaneous redox reactions into electrical energy is known as a galvanic cell or a voltaic cell.
Galvanic cell Voltaic cell is an electrochemical cell that makes use of chemical reactions to generate electrical energy

Parts of Galvanic Cell

Anode – Oxidation occurs at this electrode.
Cathode – Reduction occurs at this electrode.
Salt bridge – Contains electrolytes which are required to complete the circuit in a galvanic cell.
Half-cells – reduction and oxidation reactions are separated into compartments.
External circuit – Conducts the flow of electrons between electrodes
Load – A part of the circuit utilizes the electron to flow to perform its function.

Example of Galvanic Cell (Daniel's Cell)

Daniel's cell is an example of a galvanic cell that converts chemical energy into electrical energy. In Daniel's cell, copper ions are reduced at the cathode while zinc is oxidized at the anode.
Reactions of Daniel cell at cathode and anode are:



At cathode: $Cu^{2+} + 2e^- \rightarrow Cu$

At anode: $Zn \rightarrow Zn^{2+} + 2e^-$

The Daniell cell consists of two electrodes of dissimilar metals, Zn and Cu; each electrode is in contact with a solution of its own ion; Zinc sulphate and copper sulphate respectively.

Steps Involved in Daniell Cell

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The Daniell cell can be conventionally represented as

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

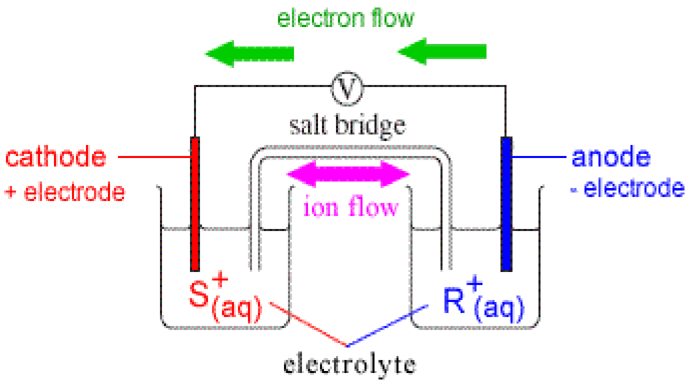
In a Daniell cell electrons flow from zinc electrode to copper electrode to copper electrode through an external circuit, while metal ions form one half cell to the other through the salt bridge.

Here current flows from copper electrode to zinc electrode that is cathode to anode via an external circuit.

Daniell cell is a reversible cell while a voltaic cell may be reversible or irreversible

Salt bridge

It maintains electrical neutrality in two compartments by allowing movement of anions towards anodic compartments and cat ions towards cathodic compartment.



Functions of Salt bridge

- A salt bridge acts as an electrical contact between the two half cells.
- It prevents mechanical flow of solution, but it provides a free path for the migration of ions, to maintain an electric current through the electrolyte solution. It prevents the accumulation of charges.
- A salt bridge helps in maintaining the charge balance in the two half cells.
- A salt bridge minimizes or eliminates the liquid junction potential.

Electrolytic Cell

An electrolytic cell is a kind of electrochemical cell. It is often used to decompose chemical compounds, in a process called electrolysis—the Greek word **lysis** means to break up. Important examples of electrolysis are the decomposition of water into hydrogen and oxygen, and bauxite into aluminium and other chemicals

An electrolytic cell has three component parts: an electrolyte and two electrodes (a cathode and an anode).

Cathode	**Anode**
Denoted by a positive sign since electrons are consumed here	Denoted by a negative sign since electrons are liberated here
A reduction reaction occurs in the cathode of an electrochemical cell	An oxidation reaction occurs here
Electrons move into the cathode	Electrons move out of the anode

Applications of Electrochemical Cells

- Electrolytic cells are used in the electro refining of many non-ferrous metals. They are also used in the electro winning of these metals.
- The production of high-purity lead, zinc, aluminum, and copper involves the use of electrolytic cells.
- Metallic sodium can be extracted from molten sodium chloride by placing it in an electrolytic cell and passing an electric current through it.
- Many commercially important batteries (such as the lead-acid battery) are made up of Galvanic cells.
- Fuel cells are an important class of electrochemical cells that serve as a source of clean energy in several remote locations.

Faraday's – First Law of Electrolysis

It is one of the primary laws of electrolysis. It states, during electrolysis, the amount of chemical reaction which occurs at any electrode under the influence of electrical energy is proportional to the quantity of electricity passed through the electrolyte.

W=ZAt

W= mass
Z= electrochemical equivalent
A= ampere current
T= time

Example 1:

A current of 0.5 ampere was passed through a solution of $CuSO_4$ for one hour. Calculate the mass of copper metal deposited on the cathode.
Electrochemical equivalent of Cu=0.000329 g/C
$= 3.29 \times 10^{-4}$ g/C or 3.294×10^{-7} Kg/C

Data:

1. Current in ampere (A) = 0.5
2. Time in second (1 hour) = 1x 60 x 60 = 3600 s
3. Z for Cu metal $= 3.294 \times 10^{-4}$ g/C $= 3.294 \times 10^{-7}$ Kg/C

Formula:

$w = Z \times A \times t$

Solution:

$w = Z \times A \times t$
$= 3.294 \times 10^{-7} \times 0.5 \times 3600$
$= 5929.2 \times 10^{-7}$ Kg
$= 5.929 \times 10^{-4}$ Kg

Answer:

Mass of copper metal deposited $= 5.929 \times 10^{-4}$ Kg or o.5929g

Faraday's – Second Law of Electrolysis

During electrolysis, when the same quantity of electricity passes through the electrolytic solution, a number of different substances liberated are proportional to their chemical equivalent weights (Equivalent weight is defined as the ratio of the atomic mass of metal and the number of electrons required for reducing the cation).
From these laws of electrolysis, we can deduce that the amount of electricity needed for oxidation-reduction depends on the stoichiometry of the electrode reaction.

What's a Faraday?

Ans: The faraday is an electric charge volume unit without measurements, equal to approximately 6.02×10^{23} electric charge carriers.

What is Electroplating?

Electroplating is basically the ***process of plating a metal onto the other by hydrolysis*** mostly to prevent corrosion of metal or for decorative purposes.

How does Electroplating Work?

To understand the concept further, let's take an example of a gold coating. In this instance, a layer of gold is to be electrodeposited on metallic jewelry to enhance its appearance.
Usually, the gold plating is connected to the anode (+ve charged electrode) of the circuit and the jewellery is kept at the cathode (-ve charged electrode).

- The voltage level of current.
- The temperature and chemical composition of the bath.
- The current length of time.
- The distance between the cathode and the anode.

Uses of Electroplating

Talking about the uses of electroplating, apart from enhancing the appearance of the substrate it is used in various other purposes as well. The major application is to optimize a material's resistance towards corrosion. Some of the other common applications of electroplating involve:

- Improving wear resistance.
- Improving the thickness of the metal surface.

Enhancing the electrica

- l conductivity like plating a copper layer on an electrical component.
- Minimizing Friction.
- Improving surface uniformity.

What is battery?

• A container consisting of one or more cells, in which chemical energy is converted into electricity and used as a source of power

Rusting of iron and its prevention

Rusting of iron refers to the formation of **rust**, a mixture of iron oxides, on the surface of iron objects or structures. This rust is formed from a redox reaction between oxygen and iron in an environment containing water (such as air containing high levels of moisture). The rusting of iron is characterized by the formation of a layer of a red, flaky substance that easily crumbles into a powder.

What is the Chemistry Behind the Rusting of Iron?

The exposure of iron (or an alloy of iron) to oxygen in the presence of moisture leads to the formation of rust. This reaction is not instantaneous, it generally proceeds over a considerably large time frame. The oxygen atoms bond with iron atoms, resulting in the formation of iron oxides. This weakens the bonds between the iron atoms in the object/structure.

- The reaction of the rusting of iron involves an increase in the oxidation state of iron, accompanied by a loss of electrons. Rust is mostly made up of two different oxides of iron that vary in the oxidation state of the iron atom. These oxides are:
- Iron (II) oxide or ferrous oxide. The oxidation state of iron in this compound is +2 and its chemical formula is FeO.
- Iron (III) oxide or ferric oxide, where the iron atom exhibits an oxidation state of +3. The chemical formula of this compound is Fe_2O_3.

Reference pages

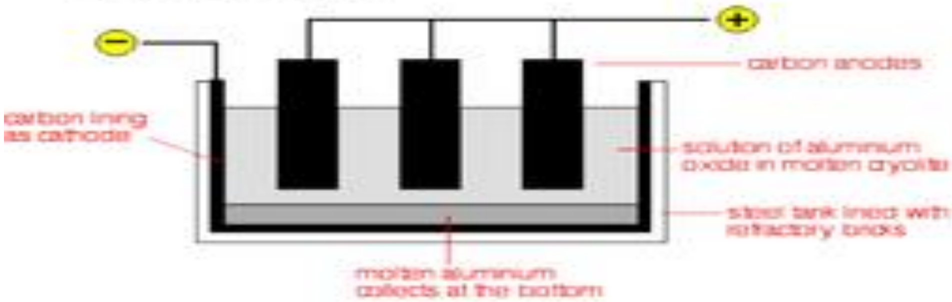
https://chem.libretexts.org/Bookshelves/Analytical_Chemistry/Supplemental_Modules_(Analytical_Chemistry)/Electrochemistry/Redox_Chemistry/Definitions_of_Oxidation_and_Reduction#:~:text=Oxidation%20and%20Reduction%20with%20respect,is%20the%20loss%20of%20oxygen.

https://www.britannica.com/science/oxidation-reduction-reaction
https://byjus.com/chemistry/electrochemical-cell/#:~:text=An%20electrochemical%20cell%20is%20a,electrical%20energy%2C%20or%20vice%20versa.
https://www.thoughtco.com/definition-of-oxidizing-agent-605459#:~:text=An%20oxidizing%20agent%20is%20a,is%20thus%20an%20electron%20acceptor.
https://byjus.com/chemistry/reducing-agent/
https://byjus.com/chemistry/daniell-cell/

Worksheets

ELECTROLYTIC CELLS WORKSHEET

1. The diagram below shows the extraction of aluminium from its molten oxide by electrolysis.



a) Name the main ore of aluminium

____

b) State why cryolite is added to the molten oxide

____

c) At what temperature is this process carried out? ____________________

d) Write half-cell equations for the reactions

At the cathode ____

At the anode ____

e) Name any other gas given off at the anode ________________________

2. The ions present in the electrolysis of aqueous copper II sulphate solution using carbon electrodes are: Cu^{2+}(aq), SO_4^{2-}(aq), H^+(aq), OH^-(aq)
   a) Write ionic half-cell equations for the reactions at the
      i) Anode

____

Lesson plan

04_Chemistry_Lesson_Plan.pdf

http://sunlight.caltech.edu/leoleary/04_Chemistry_Lesson_Plan.pdf

Chapter 08

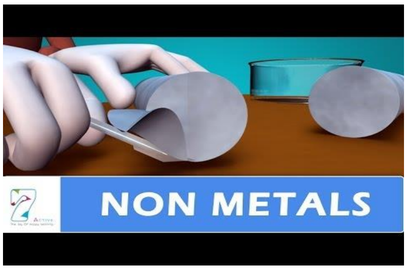
Chemical Reactivity

Chapter	Understandings	Practical
Chemical reactivity • **Metals** • Electropositive metals • Comparison of reactivity of alkali and alkaline earthy metals • Internes of Nobel metals • **Non-metals** • Comparison of reactivity of Halogens • **Essential of metals in human life** **Heavy metal test and its importance**. **NEW**	• Students will be able to: • Show how cations and anions are related to the terms metals and non-metals. • Identify elements as an alkali metal or an alkaline earth metal. (Applying) . • Explain the differences in ionization energies of alkali and alkaline earth metals. • Differentiate between soft and hard metals (Iron and Sodium). (Analyzing) • Describe the inertness of noble metals. (Understanding). • Identify the commercial value of Silver, Gold and Platinum. (Analyzing). • Compile some important reactions of halogens. (Applying)	• **How to remove rust from iron with the help of lemon.(NEW)** • **Anemic test for iron deficiency in blood.** **(NEW)**

Videos

Metals

Uses of metals



Non-Metals

Alkalain Earth metal

Alkali Metals

Halogens

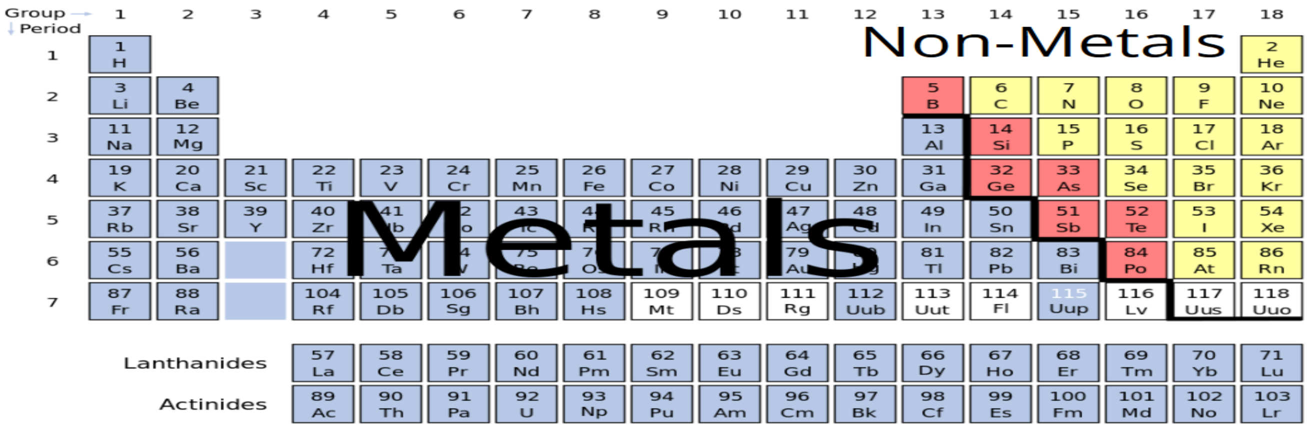
Chapter Overview

Chemical Reactivity

The chemical reactivity of a substance can refer to the variety of circumstances (conditions that include temperature, pressure, presence of catalysts) in which it reacts, in combination with the:
Variety of substances with which it reacts
Equilibrium point of the reaction (i.e., the extent to which all of it reacts)
Rate of the reaction

What is the Reactivity Series?

The reactivity series of metals, also known as the activity series, refers to the arrangement of metals in the descending order of their reactivities. The data provided by the reactivity series can be used to predict whether a metal can displace another in a single displacement reaction. It can also be used to obtain information on the reactivity of metals towards water and acids.
Metals and Nonmetals are the elements present around us. So, it is important to know whether a particular element is a metal or nonmetal. ***Materials can be divided into metals and nonmetals.*** The characteristic feature of metals like copper and aluminium is a high electrical and thermal conductivity, while nonmetals such as phosphorus metal and sulfur are insulators. Elements are distinguished as above based on their propertie**s.**



What are Metals?

Metals are minerals or substances that form naturally below the surface of the Earth. Most metals are lustrous or shiny. Metals ate inorganic, which means they are made of substances that were never alive.

Physical Properties of Metals

- All the metals are good conductors of heat and electricity. Cooking utensils and irons are made up of metals as they are good conductors of heat.
- Ductility is the ability of the material to be stretched into a wire.
- Malleability is the property of substances which allows them to be beaten into flat sheets.
- Aluminium sheets are used in the manufacturing of Aircraft because of their lightweight and strength.
- Metals are malleable.
- Metals are sonorous because it produces a deep or ringing sound when struck with another hard object.
- Usually, all the metals have a shiny appearance but these metals can also be polished to have a shiny appearance.

Chemical properties of Metals

- **Reaction with water:** Only highly reactive metals react with water and not all the metals. For example, Sodium reacts vigorously with water and oxygen and gives a large amount of heat in the process. This is why sodium is stored in kerosene so that it does not come in contact with moisture or oxygen.
- **Reaction with acids:** Hydrogen gas is produced when metals react with acids. For example, when zinc reacts with hydrochloric acid it produces zinc chloride and hydrogen gas.
- **Reaction with bases:** Not all the metals react with bases and when they do react, they produce metal salts and hydrogen gas. When zinc reacts with strong sodium hydroxide it gives sodium zincate and hydrogen gas.
- **Reaction with oxygen:** Metal oxides are produced when metals burn in the presence of oxygen. These metal oxides are basic in nature. For example: When magnesium strip is burned in the presence of oxygen it forms magnesium oxide and when magnesium oxide dissolves in water it forms magnesium hydroxide.

Before exploring the properties of metals, it's helpful to understand just what is known by the term "element" and how the periodic table is used to structure elements on the table.

Types of metals

There are also three basic types of metals:

- alkali metals,
- alkaline earth metals
- And transitional metals.

The transitional metals include a number of subcategories of their own, described later.

Periodic Trends of Alkali Metals

The alkali metals are lithium (Li), sodium (Na), potassium (K), rubidium (Rb), caesium (Cs), and francium (Fr).

This group lies in the s-block of the periodic table, as all alkali metals have their outermost electron in an s-orbital. increasing atomic radius,

decreasing electronegativity

increasing reactivity

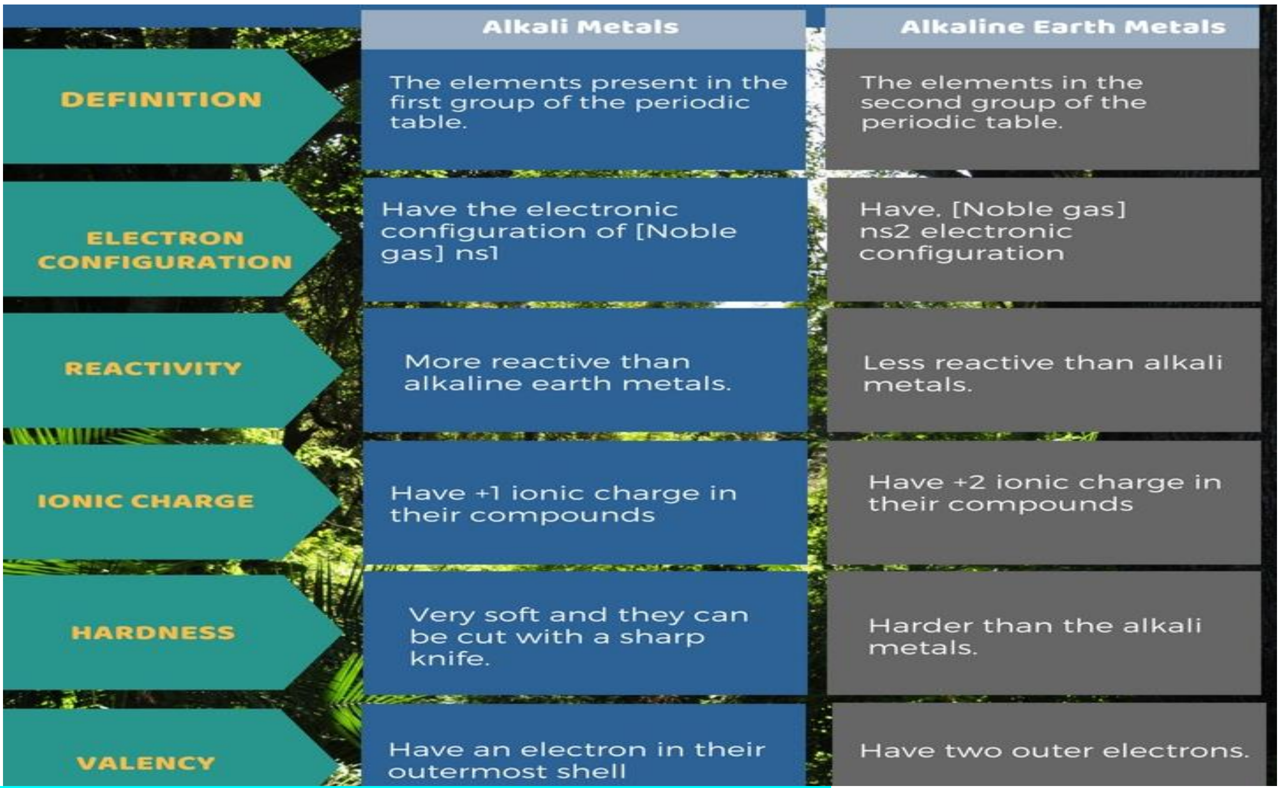
decreasing melting and boiling points

In general, their densities increase when moving down the table, with the exception of potassium, which is less dense than sodium.

Periodic Trends of Alkaline Earth Metals

The alkaline earth metals (beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra)) are a group of chemical elements in the s-block of the periodic table with very similar properties:

The alkaline earth metals are shiny, silvery-white, and somewhat reactive metals at standard temperature and pressure.
All the alkaline earth metals readily lose their two outermost electrons to form cations with a 2+ charge.
low densities
low melting points
low boiling points
The alkaline earth metals comprise the group 2 elements. All the discovered alkaline earth metals occur in nature.



	Alkali Metals	Alkaline Earth Metals
DEFINITION	The elements present in the first group of the periodic table.	The elements in the second group of the periodic table.
ELECTRON CONFIGURATION	Have the electronic configuration of [Noble gas] ns1	Have. [Noble gas] ns2 electronic configuration
REACTIVITY	More reactive than alkaline earth metals.	Less reactive than alkali metals.
IONIC CHARGE	Have +1 ionic charge in their compounds	Have +2 ionic charge in their compounds
HARDNESS	Very soft and they can be cut with a sharp knife.	Harder than the alkali metals.
VALENCY	Have an electron in their outermost shell	Have two outer electrons.

Why iron metal is hard and sodium is soft?

Why **is sodium** light and **iron hard** in nature? **Sodium** is an alkali **metal** and is very **soft** and can cut through knife while **iron** is a transition **metal** which is very **hard** and requires great energy to break apart. **Iron** is **hard**, tough and strong compared with **sodium** because of the strong metallic atom-atom bonding

Electropositive Elements

An electropositive element is one that has tendency to lose electrons and form positively charged ion. Metals like Na, Mg, K, Ca, Fe, Zn lose electrons and form positively charged ion. Due to this metals are called electropositive elements.
electropositive elements tend to lose electrons and form positive ions, e.g. the univalent alkali metals Li^+, Na^+, K^+, etc., and the divalent alkaline-earth metals Be^{2+}, Mg^{2+}, Ca^{2+}. Compare electronegativity.
Example:
Sodium metal forms positively charged sodium ion, Na^+

Noble metal,

any of several metallic chemical elements that have outstanding resistance to oxidation, even at high temperatures; the grouping is not strictly defined but usually is considered to include rhenium, ruthenium, rhodium, palladium, silver, osmium, iridium, platinum, and gold; *i.e.*, the metals of groups VIIb, VIII, and Ib of the second and third transition series of the periodic table

What are Non-Metals?

Non-metals are the elements which form negative ions by accepting or gaining electrons. Non-metals usually have 4, 5, 6 or 7 electrons in their outermost shell.

Non-metals are those which lack all the metallic attributes. They are good insulators of heat and electricity. They are mostly gases and sometimes liquid. Some they are even solid at room temperatures like Carbon, sulfur and phosphorus.

Physical Properties of Non-Metals

- Ductility is the property of the material to be stretched into wires but non-metals are not ductile except for carbon, as carbon fibres find uses in a wide variety of industries including sports and music equipment.
- Another property characteristic to metals is absent in non-metals called malleability. They can't be drawn into sheets are they are brisk and break on applying pressure.
- They are not lustrous as they do not have any shiny appearance.
- They are not sonorous and do not produce a deep ringing sound when they are hit with another material. They are also bad conductors of heat and electricity except for graphite.

Chemical Properties of Non-Metals

- **Reaction with Water**

Non-metal does not react with water but it is usually very reactive in air, which is why some of them are stored in water. For example, one of the highly reactive non-metals is phosphorus and it catches fire when exposed to air that is why it is stored in water to prevent its contact with atmospheric oxygen.

- **Reaction with Acids**

None of the non-metals is known to react with acids.

- **Reaction with Bases**

The reaction between non-metals and bases is a very complex one. The reaction of chlorine with bases like sodium hydroxide gives products like sodium hypochlorite, sodium chloride as well as water.

- **Reaction with Oxygen**

Oxides of non-metals are formed when it reacts with oxygen. The oxides of non-metals are acidic or neutral in nature.

- When sulfur reacts with oxygen, we get sulfur dioxide. $S + O_2 \rightarrow SO_2$

Uses of Non-Metals

- For the preparation of ammonia, nitric acid and fertilizers, nitrogen is used.
- For the purification of water, chlorine is used,
- Hydrogen is very useful as rocket fuel.
- Carbon can be used to make pencils when it is in the graphite form.
- Sulphuric acid is prepared using sulphur.

Halogen as non-metal

- Halogens are nonmetals in group 17 (or VII) of the periodic table. Down the group, atom size increases.
- As a diatomic molecule, fluorine has the weakest bond due to repulsion between electrons of the small atoms.
- Due to increased strength of Van der Waals forces down the group, the boiling points of halogens increase. Therefore, the physical state of the elements down the group changes from gaseous fluorine to solid iodine.
- Due to their high effective nuclear charge, halogens are highly electronegative. Therefore, they are highly reactive and can gain an electron through reaction with other elements.

Halogens can be harmful or lethal to biological organisms in sufficient quantities.

Essential metal for human.

- Several metal ions such as sodium, potassium, magnesium, and calcium are essential to sustain biological life.
- At least six additional metals, chiefly transition metals, are also essential for optimal growth, development, and reproduction, i.e. manganese, **iron**, cobalt, **copper**, **zinc**, and molybdenum.
- **Metals** like iron (Fe), zinc (Zn) and copper (Cu) are essential to **human** health.
- Cobalt, for instance, found at the core of vitamin B12, is key to making red blood cells.
- Iron allows those cells to ferry oxygen and other important chemicals to the body's tissues.

What is a heavy metal blood test?

A heavy metal blood test is a group of tests that measure the levels of potentially harmful metals in the blood. The most common metals tested for are lead, mercury, arsenic, and cadmium. Metals that are less commonly tested for include copper, zinc, aluminum, and thallium. Heavy metals are found naturally in the environment, certain foods, medicines, and even in water.

Heavy metals can get in your system in different ways. You might breathe them in, eat them, or absorb them through your skin. If too much metal gets into your body, it can cause heavy metal poisoning. Heavy metal poisoning can lead to serious health problems. These include organ damage, behavioral changes, and difficulties with thinking and memory. The specific symptoms and how it will affect you, depend on the type of metal and how much of it is in your system.

Other names: heavy metals panel, toxic metals, heavy metal toxicity test

What is it used for?

Heavy metal testing is used to find out if you have been exposed to certain metals, and how much of the metal is in your system.

Why do I need a heavy metal blood test?

Your health care provider may order a heavy metal blood test if you have symptoms of heavy metal poisoning. The symptoms depend on the type of metal and how much exposure there was.

Your symptoms may include:

- Nausea, vomiting, and abdominal pain
- Diarrhea
- Tingling in the hands and feet
- Shortness of breath and Chills
- Weakness

Reference pages

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lesson plan

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worksheet

Choose the correct option:

1. The most reactive metal is:

(a) Iron (b) Gold (c) Zinc (d) Potassium.

2. The liquid metal at room temperature

(a) Mercury (b) Bromine (c) Sodium (d) Gold.

3. Non-metals are:

(a) generally liquids (b) generally gases (c) generally solids and gases (d) generally gases and liquids.

4. **The metal which is stored in kerosene:**

(a) Phosphorus
(b) Magnesium
(c) Sodium
(b) Magnesium

5. **The non-metal which is liquid at room temperature is:**

(a) Carbon
(b) Iodine
(c) Bromine
(d) Chlorine.

6. Materials around us can be classified into

(a) Elements and compounds
(b) Metals and non-metals
(c) Acids and bases
(d) None of these.

7. All metals are solids except

(a) Sodium
(b) Calcium
(c) Mercury
(d) Hydrogen.

8. Metal oxides are of nature

(a) Acidic
(b) Basic
(c) Neutral
(d) All of these.

9. The metal which can be cut with a knife

a) Sodium and potassium
b) Barium and calcium
c) Sodium and mercury
d) Potassium and calcium.

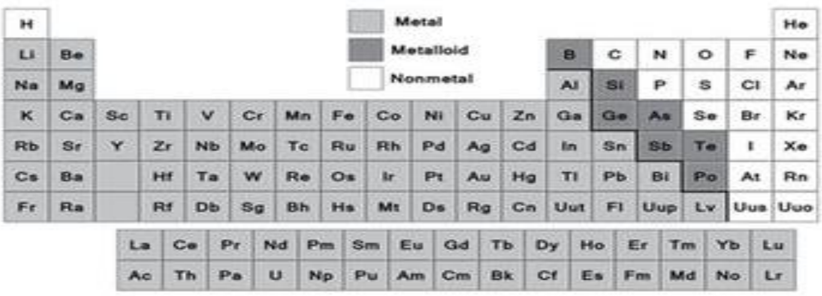
10. When non-metals react with water then

(a) Hydrogen gas is formed
(b) Carbon dioxide gas is formed
(c) Non-metals do not react with water
(d) None of these.

Answers

1. (d)	2. (a)	3. (c)	4. (c)	5. (c)
6. (b)	7. (c)	8. (b)	9. (a)	10. (c)

Metals , Nonmetals and Metalloids



Identify each element as a metal, nonmetal or metalloid.

Copper	
Oxygen	
Boron	
Potassium	
Silicon	
Helium	
Aluminum	
Hydrogen	
Calcium	
Polonium	

List four physical properties of metals, nonmetals and metalloids.

Metals	Nonmetals	Metalloids

What's New in it?

Although it is based on National curriculum and Sindh curriculum.

New Practical's	Topic
Grow crystal of sugar (NEW)	• Essential of metals in human life. • Heavy metal test and its importance. NEW
Preparation of alum crystal. (NEW)	
How to remove rust from iron with the help of lemon. (NEW)	
Anemic test for iron deficiency in blood. (NEW)	
Prepare the pure copper sulphate crystal from given impure sample. (NEW)	

Education opens up the mind, expands it and allows you to improve your life in so many ways.

If you want to inspire and make change in the lives of your students than I hope you'll find this material helpful.

Thankyou