

UNIT 4

KEY CONTENT AREA OF SCIENCE-

UNDERSTANDING MATTER

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- Structure of atom and molecules, isotopes and isobars
- Periodic properties of elements (link to atomic structure)
- Types of chemical bonding and chemical reactions
- Element - their extraction, purification of organic compound
- Acids, bases and salts
- Carbon
- Organic compounds classification and nomenclature of organic compound, isomerism organic compound
- Relationship between functioning of different organ systems
- Ecosystem - structure and function
- Evolution - Theories of Evolution, Human Evolution, diversity - level of Biodiversity, uses and threats of biodiversity, heredity, Ecological adaptations.

Structure of atom and molecules, isotopes and isobars

LEARNING OBJECTIVES

Discuss the electronic and structural properties of an atom

LEARNING OBJECTIVES

Discuss the properties of isotopes and their use in radiometric dating

KEY TAKEAWAYS

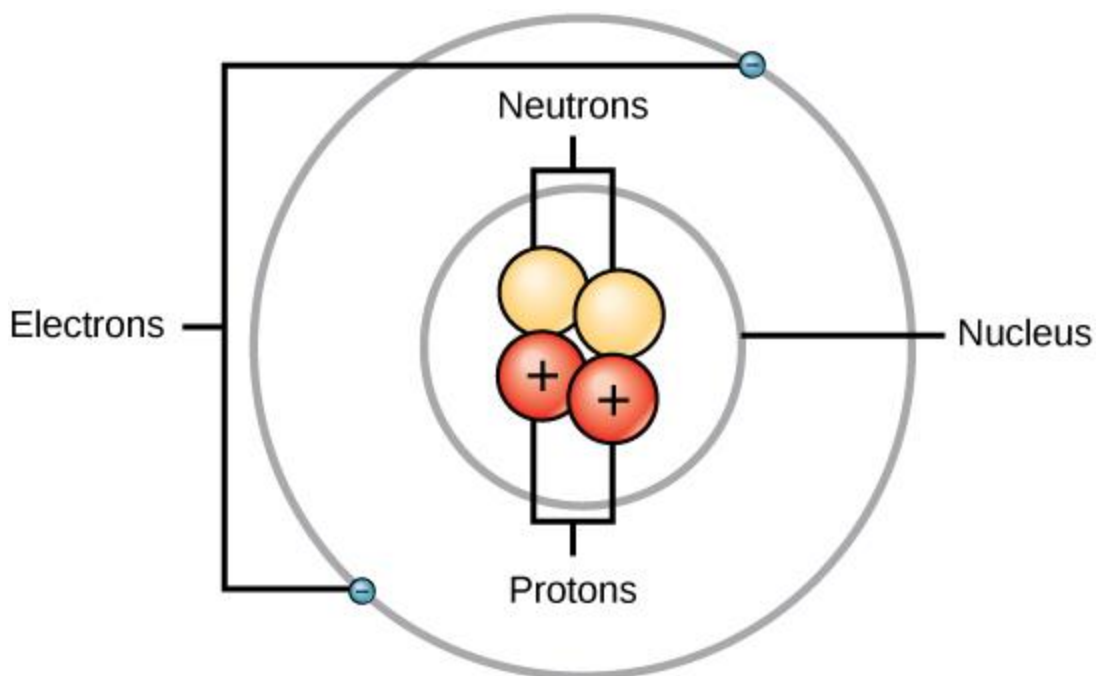
Key Points

- An atom is composed of two regions: the nucleus, which is in the center of the atom and contains protons and neutrons, and the outer region of the atom, which holds its electrons in orbit around the nucleus.
- Protons and neutrons have approximately the same mass, about 1.67×10^{-24} grams, which scientists define as one atomic mass unit (amu) or one Dalton.
- Each electron has a negative charge (-1) equal to the positive charge of a proton (+1).
- Neutrons are uncharged particles found within the nucleus.

Key Terms

- **atom**: The smallest possible amount of matter which still retains its identity as a chemical element, consisting of a nucleus surrounded by electrons.
- **proton**: Positively charged subatomic particle forming part of the nucleus of an atom and determining the atomic number of an element. It weighs 1 amu.
- **neutron**: A subatomic particle forming part of the nucleus of an atom. It has no charge. It is equal in mass to a proton or it weighs 1 amu.

- An atom is the smallest unit of matter that retains all of the chemical properties of an element. Atoms combine to form molecules, which then interact to form solids, gases, or liquids. For example, water is composed of hydrogen and oxygen atoms that have combined to form water molecules. Many biological processes are devoted to breaking down molecules into their component atoms so they can be reassembled into a more useful molecule.

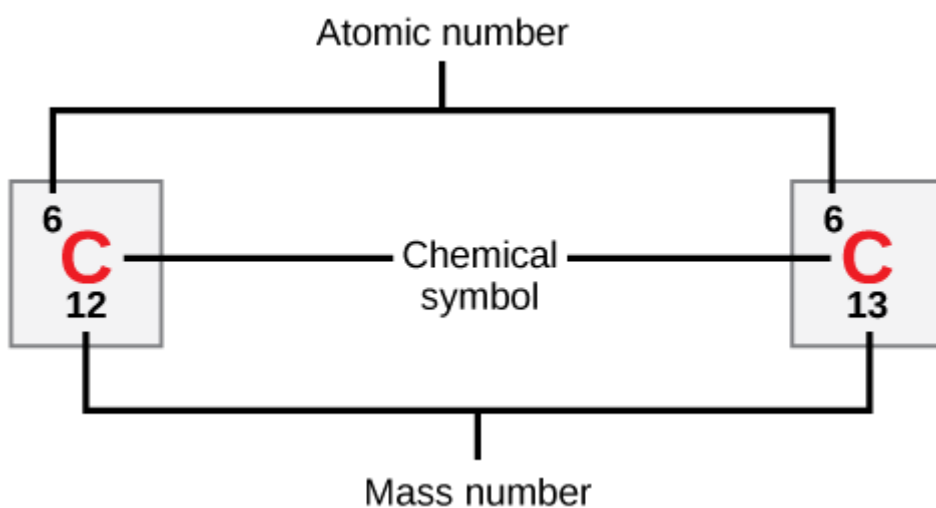


Atomic Particles

Atoms consist of three basic particles: protons, electrons, and neutrons. The nucleus (center) of the atom contains the protons (positively charged) and the neutrons (no charge). The outermost regions of the atom are called electron shells and contain the electrons (negatively charged). Atoms have different properties based on the arrangement and number of their basic particles.

The hydrogen atom (H) contains only one proton, one electron, and no neutrons. This can be determined using the atomic number and the mass number of the element (see the concept on atomic numbers and mass numbers).

Protons, Neutrons, and Electrons			
	Charge	Mass (amu)	Location
Proton	+1	1	nucleus
Neutron	0	1	nucleus
Electron	-1	0	orbitals



What is an Isotope?

Isotopes are various forms of an element that have the same number of protons but a different number of neutrons. Some elements, such as carbon, potassium, and uranium, have multiple naturally-occurring isotopes. Isotopes are defined first by their element and then by the sum of the protons and neutrons present.

- Carbon-12 (or ^{12}C) contains six protons, six neutrons, and six electrons; therefore, it has a mass number of 12 amu (six protons and six neutrons).

- Carbon-14 (or ^{14}C) contains six protons, eight neutrons, and six electrons; its atomic mass is 14 amu (six protons and eight neutrons).

While the mass of individual isotopes is different, their physical and chemical properties remain mostly unchanged.

Isotopes do differ in their stability. Carbon-12 (^{12}C) is the most abundant of the carbon isotopes, accounting for 98.89% of carbon on Earth. Carbon-14 (^{14}C) is unstable and only occurs in trace amounts. Unstable isotopes most commonly emit alpha particles (He^{2+}) and electrons. Neutrons, protons, and positrons can also be emitted and electrons can be captured to attain a more stable atomic configuration (lower level of potential energy) through a process called radioactive decay.

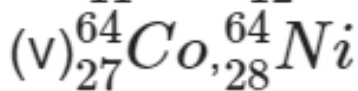
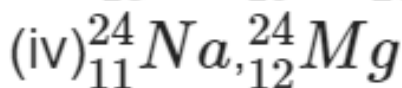
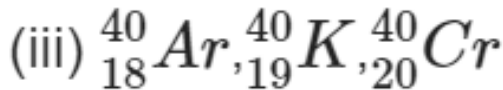
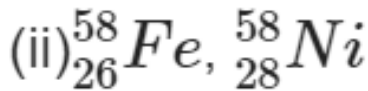
Isobars

- Species or atoms having same atomic mass number but different atomic number
- For example - ${}^{14}_6\text{C}$, ${}^{14}_7\text{N}$
- They have the same number of nucleons i.e sum of protons and neutrons are same but Number of protons and neutrons alone varies between them

Isotope	Protons(P)	Neutrons(N)	Sum(P+N)	Electrons
${}^6\text{C}^{14}$	6	8	14	6
${}^7\text{N}^{14}$	7	7	14	7

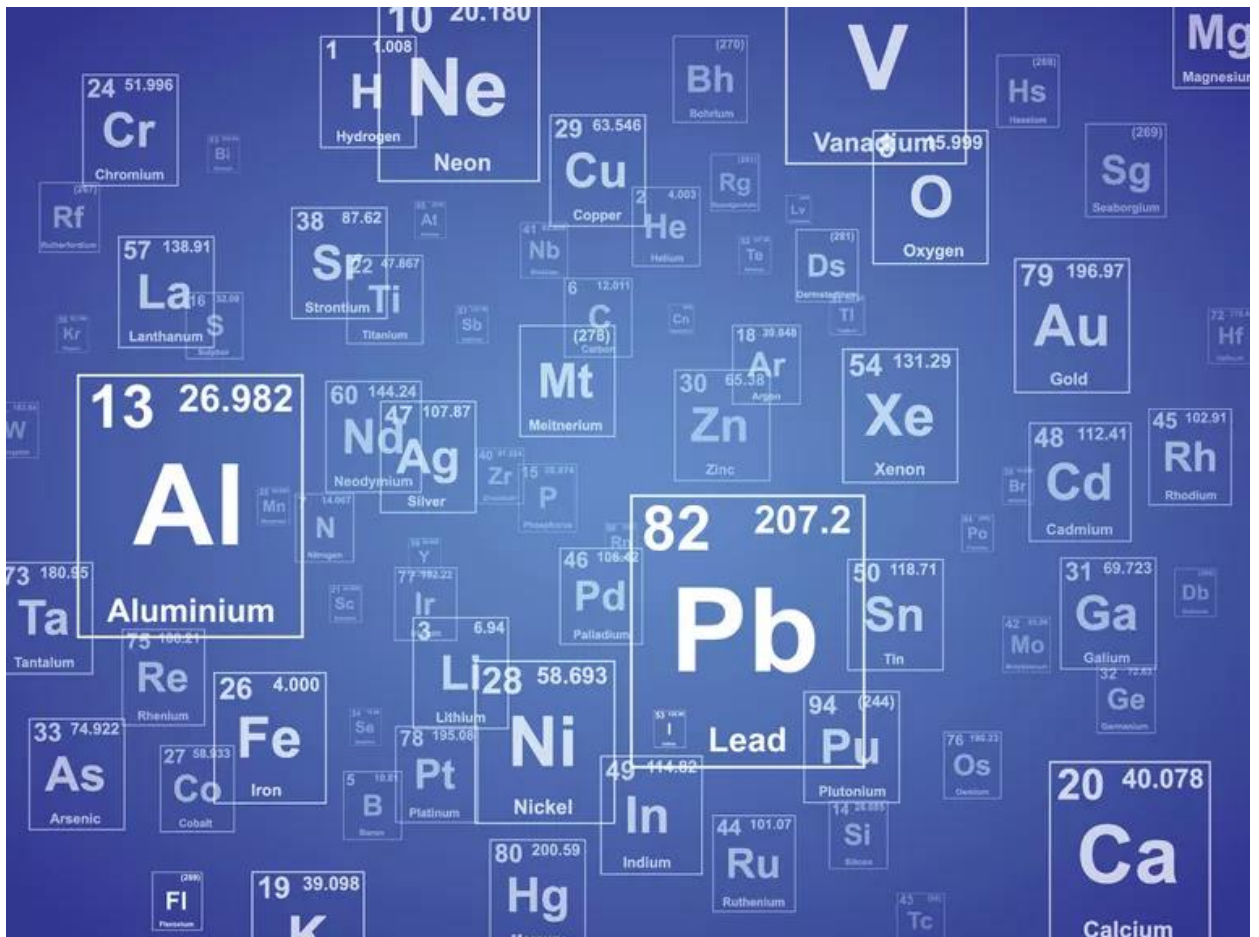
- Chemical properties depends on Atomic Number, So isobars are different chemical elements and have different chemical properties

Examples of Isobars:



Isotopes	Isobars
Isotopes are different atomic structure of same elements having same atomic number but different atomic mass number	Isobars are different elements having same Atomic mass number but different atomic number
Atomic Number is same	Atomic Number are different
Chemical properties are same	Chemical properties are different as different element
Physical properties are different	Physical properties can be same
Number of protons & electrons are same. Neutrons only differ	All neutrons, protons, and electrons differ

Periodic Properties of Elements



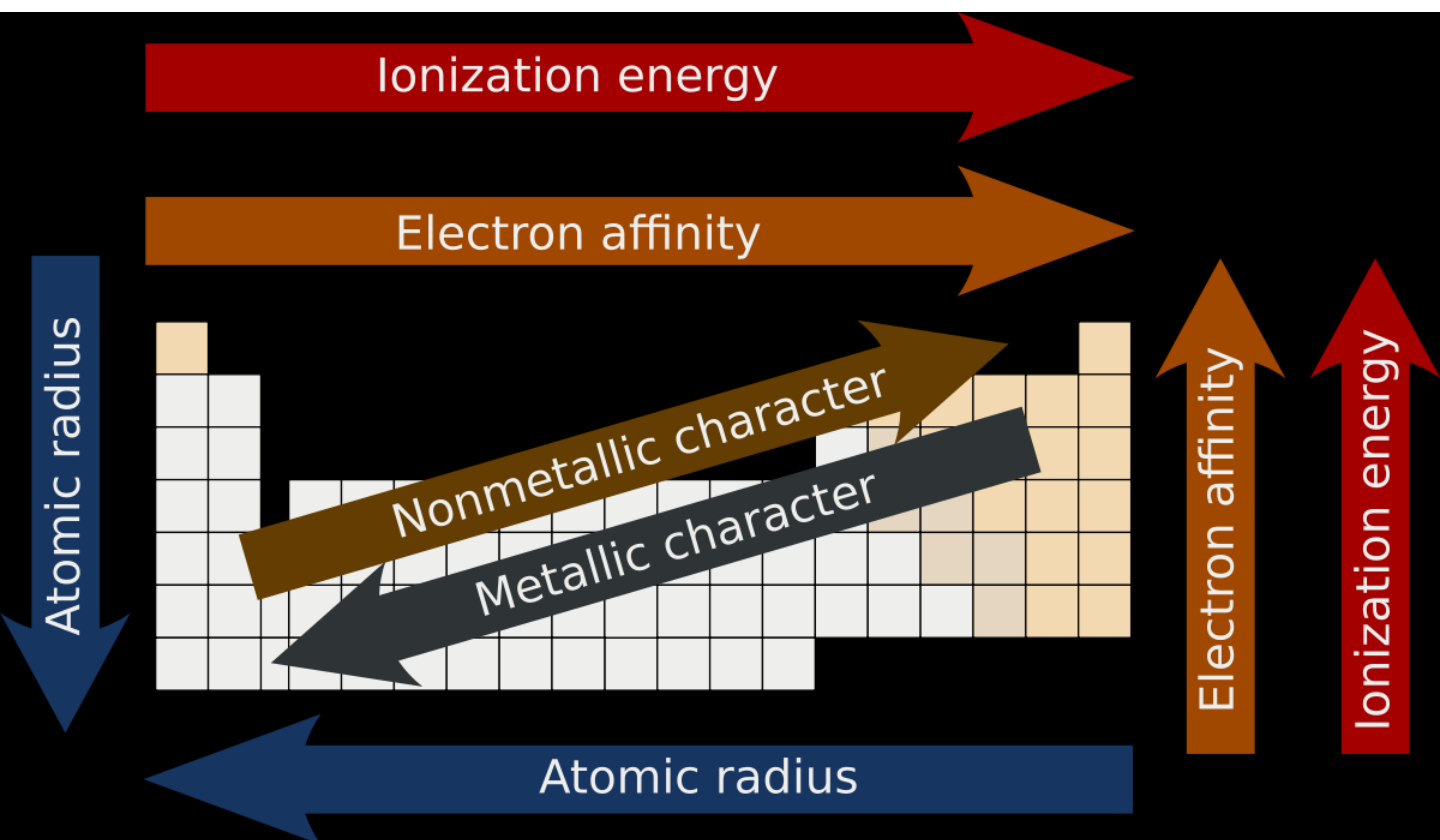
1 H Hydrogen																	2 He Helium
3 Li Lithium	4 Be Beryllium											5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
11 Na Sodium	12 Mg Magnesium											13 Al Aluminum	14 Si Silicon	15 P Phosphorus	16 S Sulfur	17 Cl Chlorine	18 Ar Argon
19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton
37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Tc Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon
55 Cs Cesium	56 Ba Barium	57-71 Lanthanides	72 Hf Hafnium	73 Ta Tantalum	74 W Tungsten	75 Re Rhenium	76 Os Osmium	77 Ir Iridium	78 Pt Platinum	79 Au Gold	80 Hg Mercury	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radon
87 Fr Francium	88 Ra Radium	89-103 Actinides	104 Rf Rutherfordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110 Ds Darmstadtium	111 Rg Roentgenium	112 Cn Copernicium	113 Nh Nihonium	114 Fl Flerovium	115 Mc Moscovium	116 Lv Livermorium	117 Ts Tennessine	118 Og Oganesson

57 La Lanthanum	58 Ce Cerium	59 Pr Praseodymium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolinium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium	71 Lu Lutetium
89 Ac Actinium	90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium	103 Lr Lawrencium

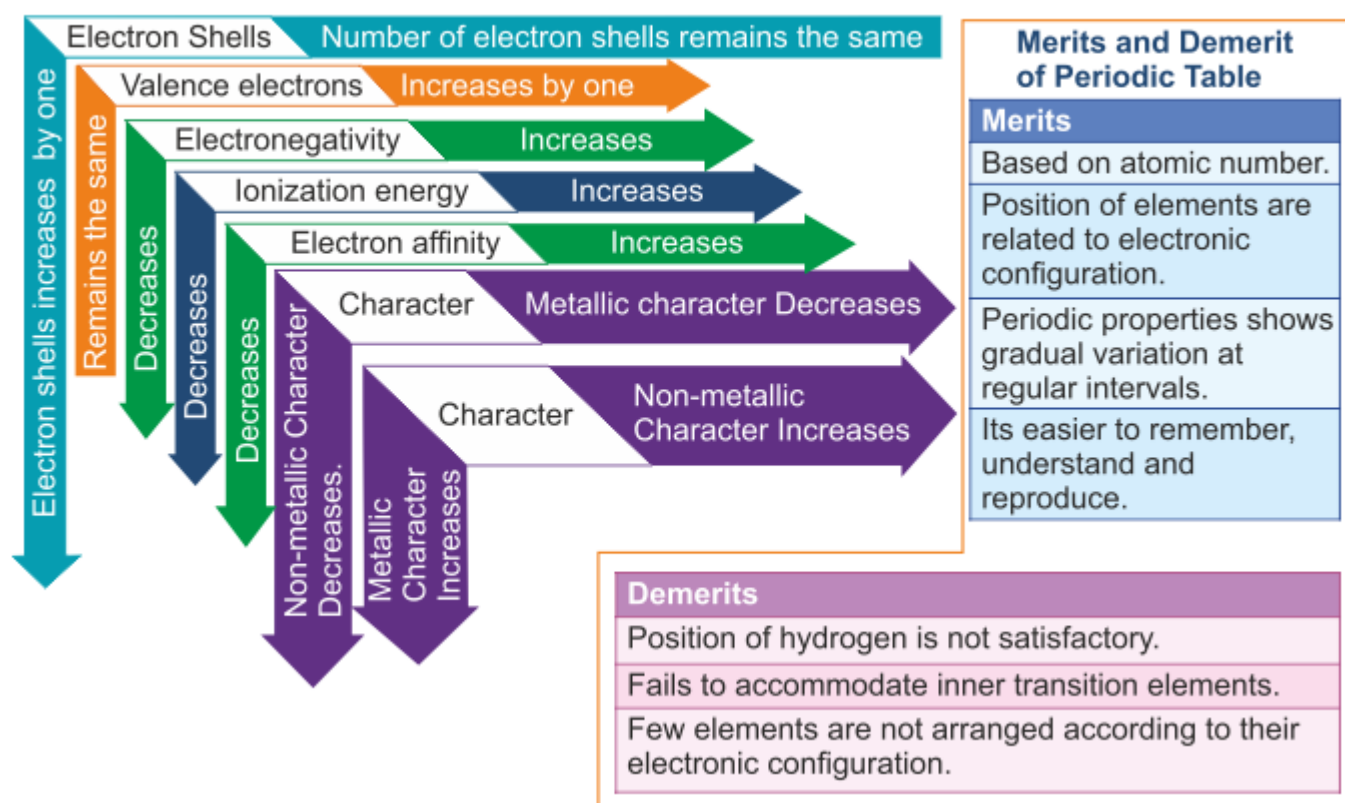
The periodic table arranges the elements by periodic properties, which are recurring trends in physical and chemical characteristics. These trends can be predicted merely by examining the periodic table and can be explained and understood by analyzing the electron configurations of the elements.

Elements tend to gain or lose valence electrons to achieve stable octet formation. Stable octets are seen in the inert gases, or noble gases, of Group VIII of the periodic table. In addition to this activity, there are two other important trends. First, electrons are added one at a time moving from left to right across a period.

As this happens, the electrons of the outermost shell experience increasingly strong nuclear attraction, so the electrons become closer to the nucleus and more tightly bound to it. Second, moving down a column in the periodic table, the outermost electrons become less tightly bound to the nucleus. This happens because the number of filled principal energy levels (which shield the outermost electrons from attraction to the nucleus) increases downward within each group.



Periodic Properties



Merits and Demerit of Periodic Table

Merits

- Based on atomic number.
- Position of elements are related to electronic configuration.
- Periodic properties shows gradual variation at regular intervals.
- Its easier to remember, understand and reproduce.

Demerits

- Position of hydrogen is not satisfactory.
- Fails to accommodate inner transition elements.
- Few elements are not arranged according to their electronic configuration.

Properties	Across the period (left to right)	Down the group (top to bottom)
Atomic size	Decreases	Increases
No. of valence electrons	Increases	Remains same
Metallic character	Decreases	Increases
Non-metallic character	Increases	Decreases
Electron affinity	Increases	Decreases
Electronegativity	Increases	Decreases
Ionisation energy	Increases	Decreases
Basic nature of oxides	Decreases	Increases
Melting point	Increases from Group I to Group IV and then decreases	Decreases in Groups I and II Decreases in Groups III and IV Increases in Groups V to VII
Boiling point	Increases from Group I to Group IV and then decreases	Decreases in Groups I and II Decreases in Groups III and IV Increases in Groups V to VII
Oxidising nature	Increases	Decreases
Reducing nature	Decreases	Increases

Types of chemical bonding and chemical reactions

Why form chemical bonds? The basic answer is that atoms are trying to reach the most stable (lowest-energy) state that they can. Many atoms become stable when their valence shell is filled with electrons or when they satisfy the octet rule (by having eight valence electrons). If atoms don't have this arrangement, they'll "want" to reach it by gaining, losing, or sharing electrons via bonds.

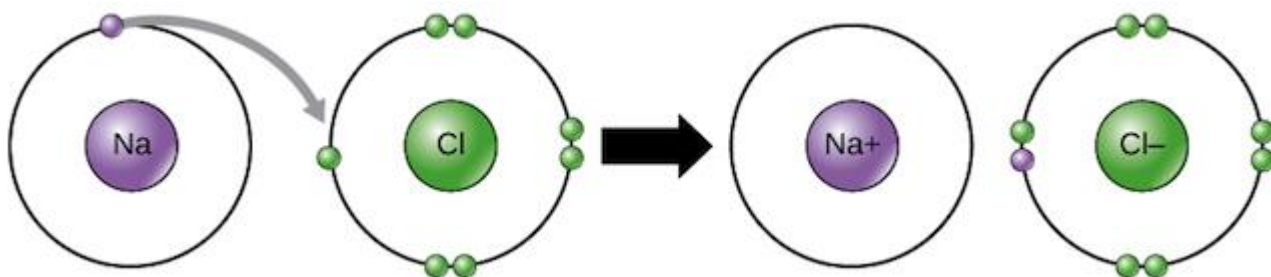
Ions and ionic bonds

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. Electron gain or loss can give an atom a filled outermost electron shell and make it energetically more stable.

Forming ions

Ions come in two types. **Cations** are positive ions formed by losing electrons. For instance, a sodium atom loses an electron to become a sodium cation. Negative ions are formed by electron gain and are called **anions**. Anions are named using the ending "-ide": for example, the anion of chlorine is called chloride.

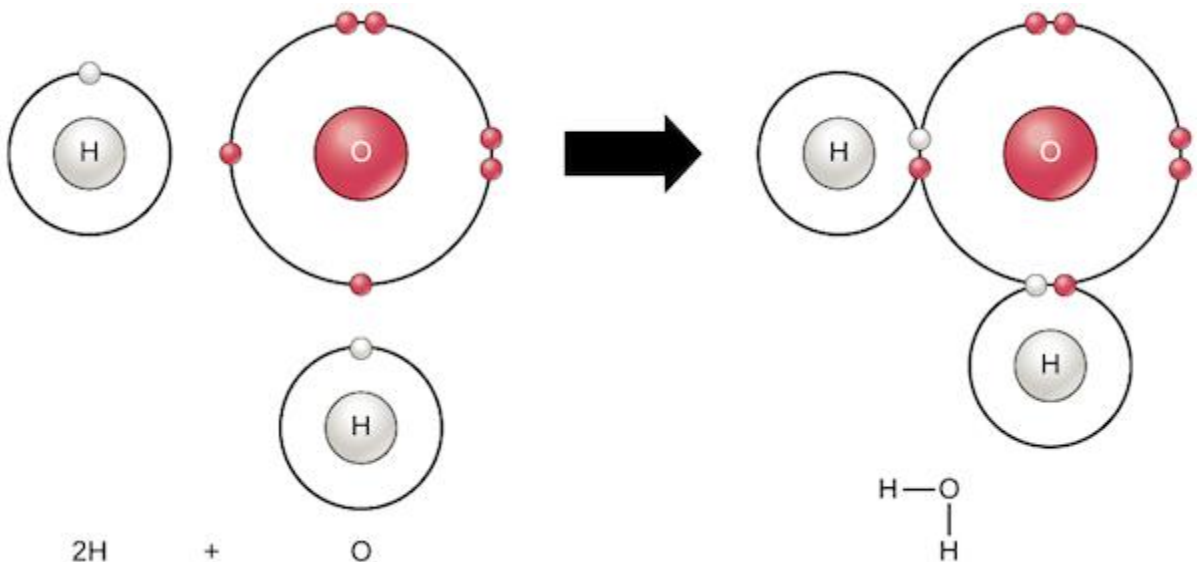
When one atom loses an electron and another atom gains that electron, the process is called **electron transfer**. Sodium and chlorine atoms provide a good example of electron transfer.



Covalent bonds

Another way atoms can become more stable is by sharing electrons (rather than fully gaining or losing them), thus forming **covalent bonds**. Covalent bonds are more common than ionic bonds in the molecules of living organisms.

For instance, covalent bonds are key to the structure of carbon-based organic molecules like our DNA and proteins. Covalent bonds are also found in smaller inorganic molecules, such as H₂O, CO₂, O₂. One, two, or three pairs of electrons may be shared between atoms, resulting in single, double, or triple bonds, respectively. The more electrons that are shared between two atoms, the stronger their bond will be.



The shared electrons split their time between the valence shells of the hydrogen and oxygen atoms, giving each atom something resembling a complete valence shell (two electrons for H, eight for O). This makes a water molecule much more stable than its component atoms would have been on their own.

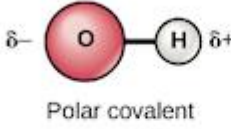
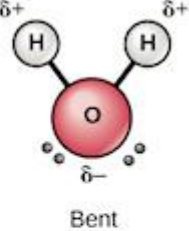

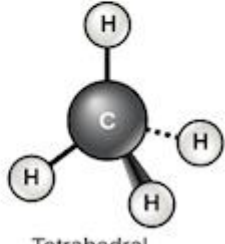
Polar covalent bonds

There are two basic types of covalent bonds: polar and nonpolar. In a **polar covalent bond**, the electrons are unequally shared by the atoms and spend more time close to one atom than the other. Because of the unequal distribution of electrons between the atoms of different elements, slightly positive (δ^+) and slightly negative (δ^-) charges develop in different parts of the molecule.

In a water molecule (above), the bond connecting the oxygen to each hydrogen is a polar bond. Oxygen is a much more **electronegative** atom than hydrogen, meaning that it attracts shared electrons more strongly, so the oxygen of water bears a partial negative charge (has high electron density), while the hydrogens bear partial positive charges (have low electron density).

Nonpolar covalent bonds

Nonpolar covalent bonds form between two atoms of the same element, or between atoms of different elements that share electrons more or less equally. For example, molecular oxygen O_2 is nonpolar because the electrons are equally shared between the two oxygen atoms.

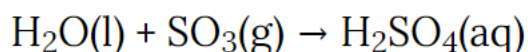
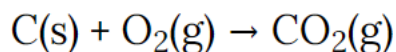
	Bond type	Molecular shape
Water	 <p>Polar covalent</p>	 <p>Bent</p>
Methane	 <p>Nonpolar covalent</p>	 <p>Tetrahedral</p>

Types of Chemical Reactions

1. Synthesis reactions

Two or more reactants combine to make 1 new product.

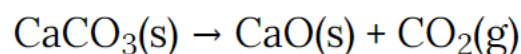
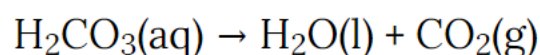
Examples:



2. Decomposition reactions

A single reactant breaks down to form 2 or more products.

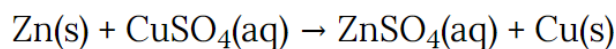
Examples:



3. Single-replacement reactions

A single element replaces a similar element of an adjacent reactant compound.

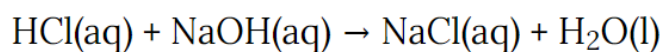
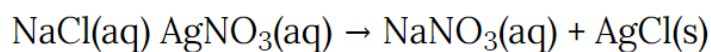
Examples:



4. Double-replacement reactions

Two ionic compounds exchange ions, producing 2 new ionic compounds.

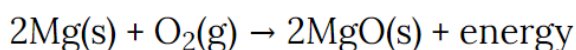
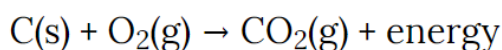
Examples:



5. Combustion reactions

A single element or compound combines with oxygen gas releasing energy. This rapid oxidation is called burning.

Examples:

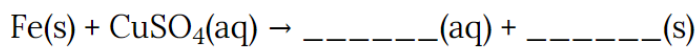


Single-replacement activity

1. Hold the test tube containing 0.5 M CuSO₄ solution at a 45° angle and insert the nail.

2. Place the test tube in the test tube rack. What happens?

3. If you use a plain iron nail, here is the reaction:



A galvanized nail (coated with Zn), undergoes a different reaction:



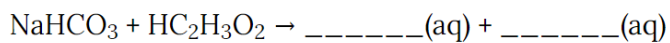
4. Complete and balance both equations.

Double-replacement activity

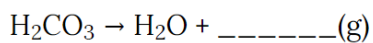
1. Pour the baking soda (sodium hydrogen carbonate, NaHCO₃) from the spoon into the 250-mL beaker containing the vinegar (acetic acid, HC₂H₃O₂).

2. Describe what happens.

3. Complete and balance the equation below for this reaction:



4. One of the products, carbonic acid (H₂CO₃), immediately decomposes into water and a gas. Complete and balance this equation, and identify the gas with a flaming or glowing splint:



Element - their extraction, purification of organic compound

Types of Purification

A large number of methods are available for the purification of substances. The choice of method, however, depends upon the nature of substance (whether solid or liquid). It also depends on the type of impurities present in it. Commonly used these methods for purification of substances:


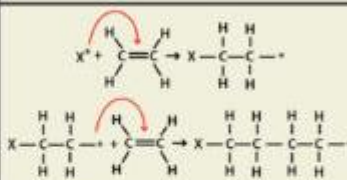
- Simple crystallisation
- Fractional crystallisation
- Sublimation
- Simple distillation
- Fractional distillation
- Distillation under reduced pressure
- Steam distillation
- Azeotropic distillation
- Chromatography

Simple Crystallisation

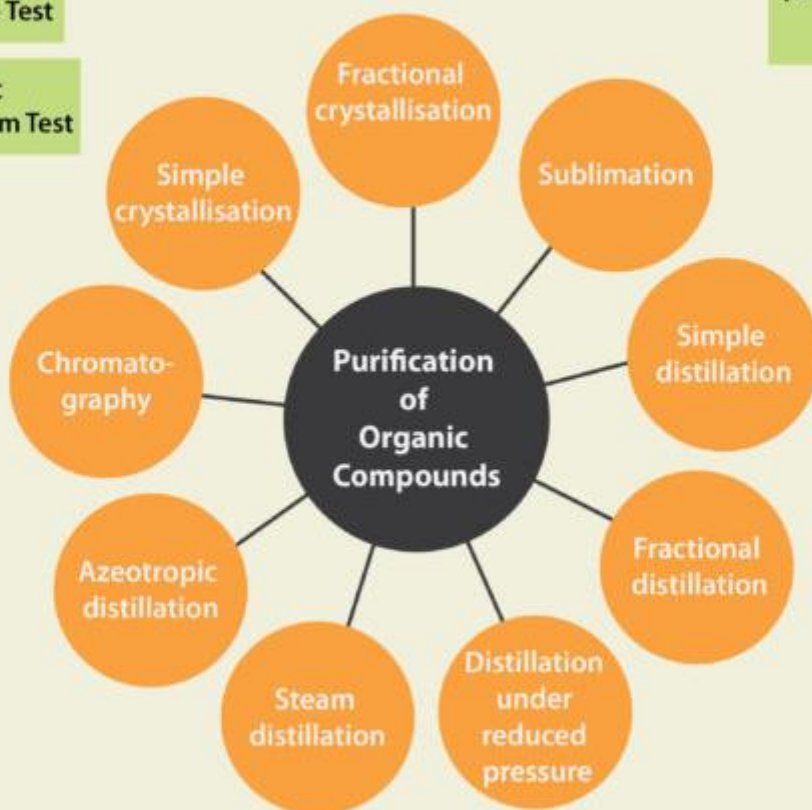
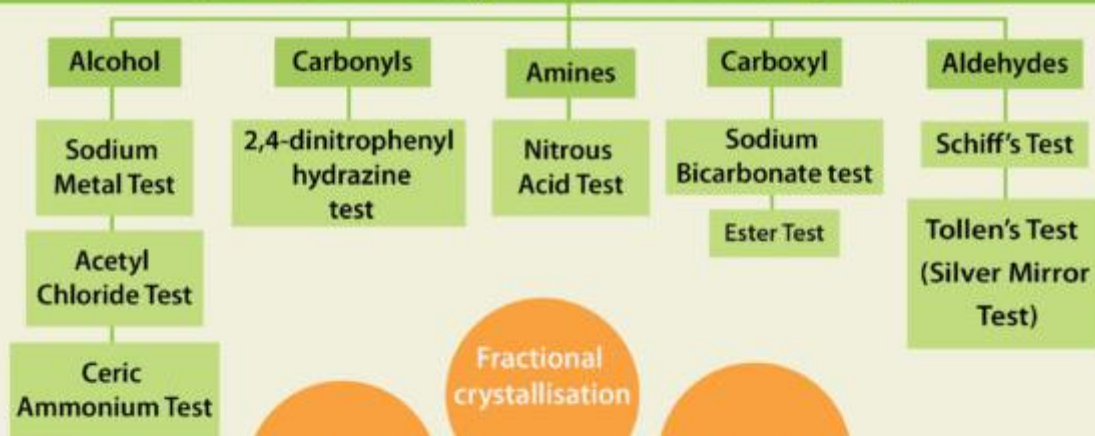
This is the most common method that we use to purify organic solids. For crystallisation, a suitable solvent is one:

- which dissolves more of the substance at a higher temperature than at room temperature
- in which impurities are either insoluble or dissolve to an extent that they remain in solution (in the mother liquor) upon crystallization
- which is not highly inflammable
- which does not react chemically with the compound to be crystallized. The most commonly-used solvents for crystallisation are water, alcohol, ether, chloroform, carbon-tetrachloride, acetone, benzene, petroleum ether etc.

Types of Organic Reactions

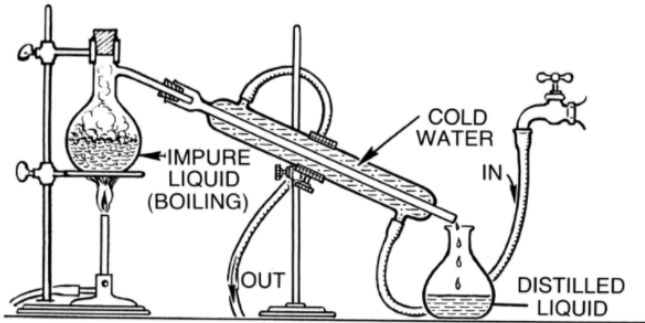
Reagent / Reaction Type	Nucleophilic	Electrophilic	Radical
Substitution $R-X + Y \rightarrow R-Y + X$	Alkyl halides + Cyanide ion or hydroxide ion $R-Br + CN^- \rightarrow R-CN + Br^-$	Benzene electrophilic substitution 	Halogenation of methane UV light $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$
Addition $C=C + XY \rightarrow C(X)C(Y)$	$C=O + H^+ \rightarrow C(OH)H$	Alkene + HBr $C=C + H-Br \rightarrow C(H)C(Br)$	
Elimination $C(X)C(Y) \rightarrow C=C + XY$	$C(Br)C(OH) \xrightarrow{OH^-} C=C + Br^- + H_2O$	Alcohol + H ⁺ acid catalysed elimination $C(OH)C(H) \xrightarrow{H^+} C=C + H_2O$	Thermal cracking of alkanes heat decane, C ₁₀ H ₂₂ → octane, C ₈ H ₁₈ + ethene, C ₂ H ₄

Qualitative Analysis of functional groups



Simple Distillation

Distillation is the joint process of vapourisation and condensation. We use this method for the purification of liquids which boil without decomposition and contain non-volatile impurities. We can also use this method for separating liquids having sufficient difference in their boiling points.



Solved Examples for You

Question: Give two practical applications of simple crystallisation.

Answer: Practical applications of simple crystallisation include:

- Sugar having an impurity of common salt can be crystallized from hot ethanol since sugar dissolves in hot ethanol but common salt does not.
- A mixture of benzoic acid and naphthalene can be separated from hot water in which benzoic acid dissolves but naphthalene does not.

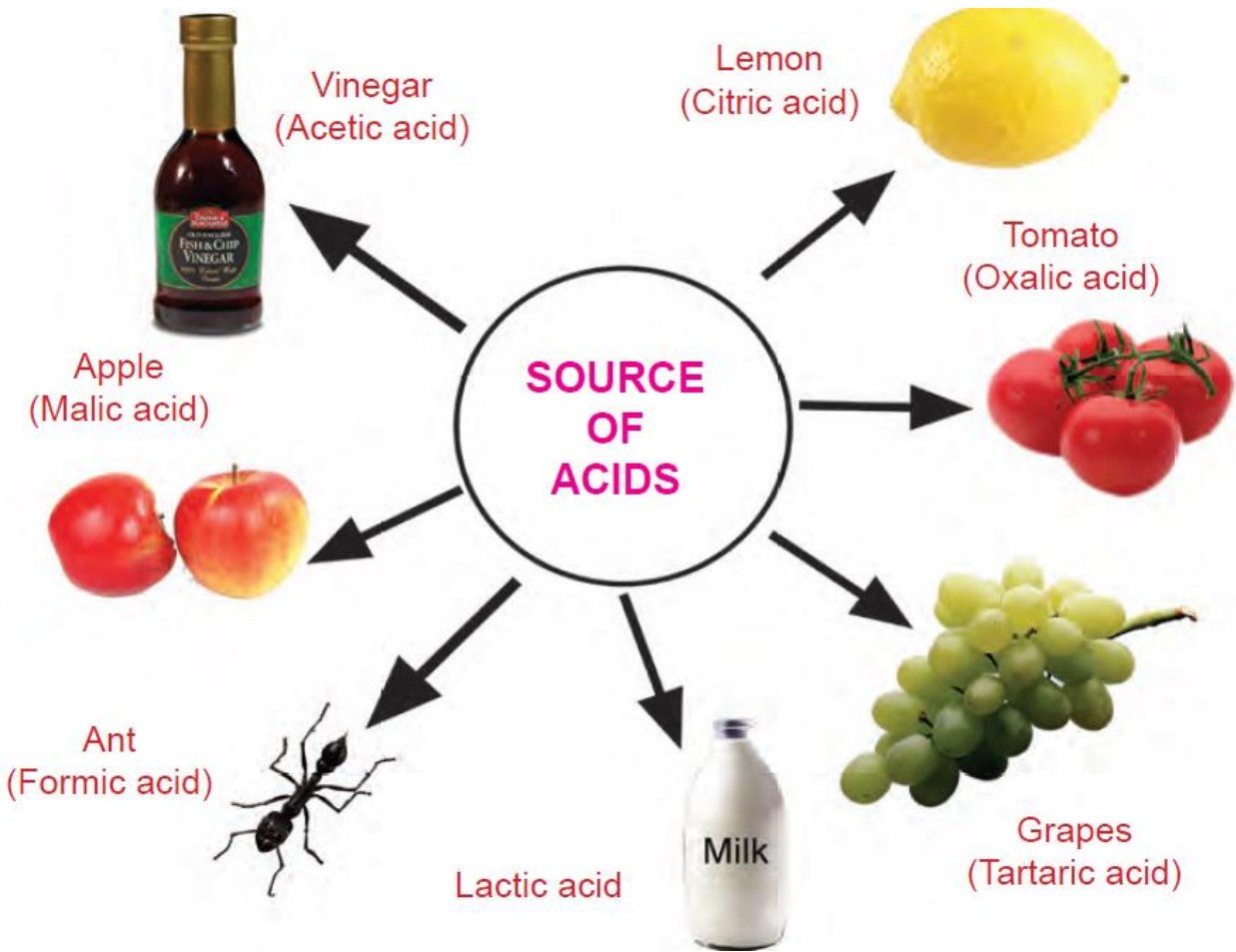
Chromatography: It is a technique used to separate mixtures into their components, purify. Compounds which can be solid or liquid and also test the purity of compounds.

- In this technique the mixture of substances are applied on to a stationary phase. Then a pure solvent or a mixture of solvents is allowed to move slowly over stationary phase. The components of mixture get gradually separated from one another. The moving phase is mobile phase.

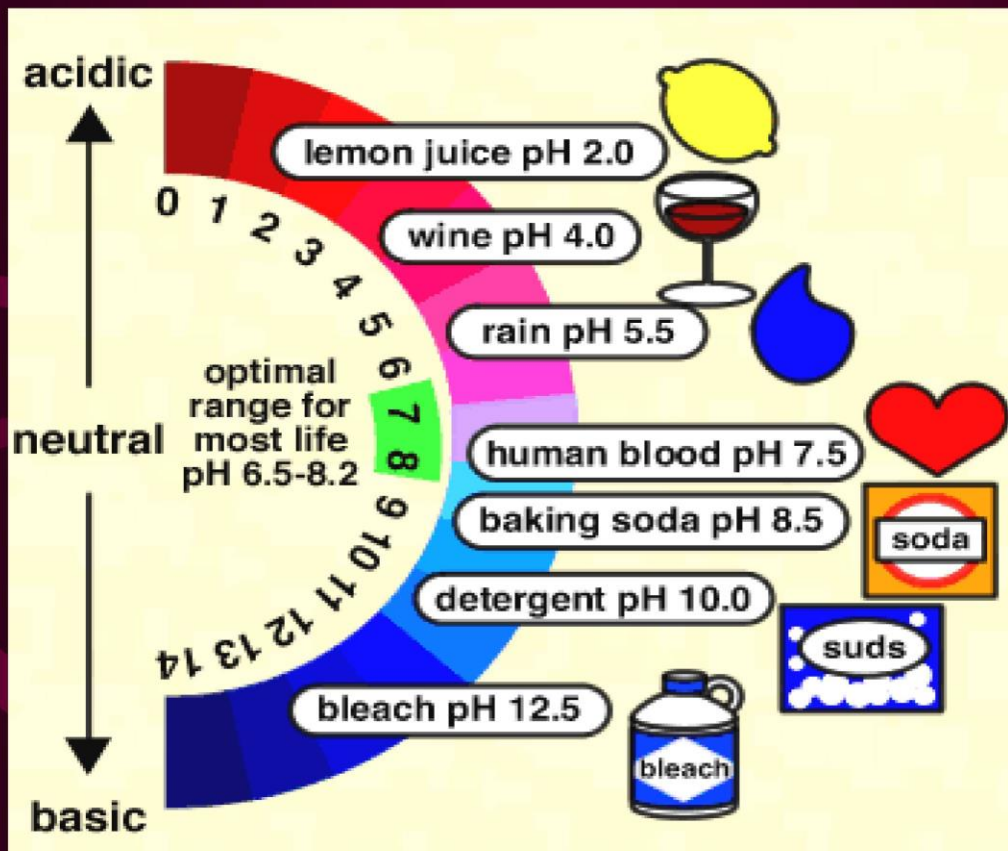
The chromatography is classified into two categories:

- Adsorption chromatography
- Partition chromatography

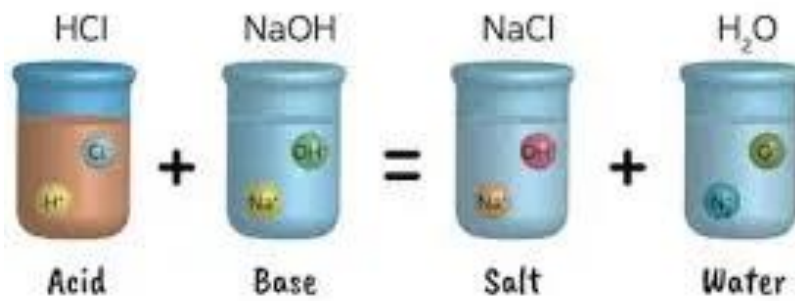
Acids, bases and salts



pH Scale



Acid - Base reactions



Acids

Acidity is a characteristic property of acids. Acidic substances are usually very sour. Apart from hydrochloric acid, there are many other types of acids around us. Citrus fruits like lemons and oranges contain citric and ascorbic acids while tamarind paste contains tartaric acid.

In fact, the word 'acid' and 'acidity' are derived from the Latin word 'acidus' which means sour. If you dip a blue litmus paper into an acid, it will turn red while a red litmus paper will not change colour. Acids also liberate dihydrogen when they react with some [metals](#).

Bases

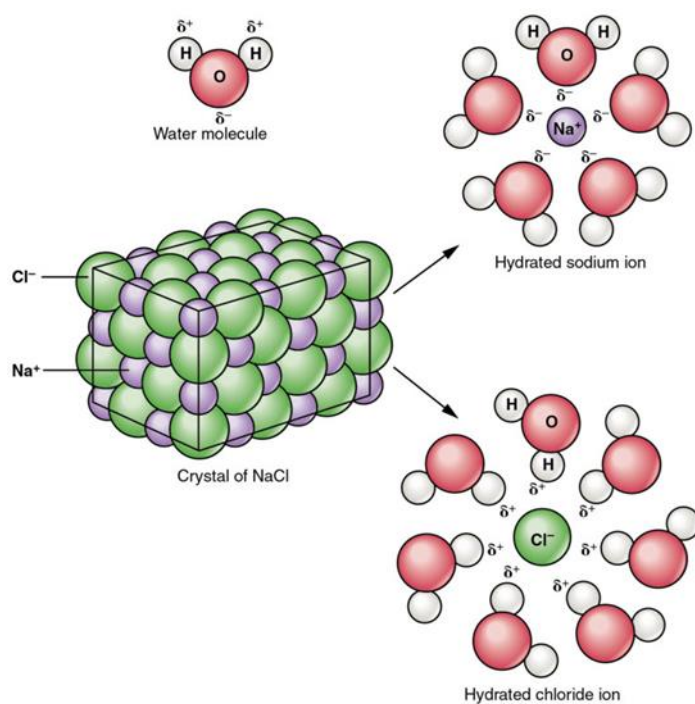
Bases turn red litmus paper blue while the blue litmus paper stays blue. They taste bitter and also feel soapy. Some other common examples of bases include sodium bicarbonate that is used in cooking and household bleach.

Salts

Apart from sodium chloride, other common salts are sodium nitrate, barium sulfate etc. Sodium chloride or common salt is a product of the reaction between the hydrochloric acid (acid) and sodium hydroxide (base). Solid sodium chloride is made of a cluster of positively charged sodium ions and negatively charged chloride ions held together by electrostatic forces.

Electrostatic **forces** between opposite charges are inversely proportional to the dielectric constant of the medium. In other words, we can say that a compound that has acidity in its nature and a compound that has basicity as its nature, may yield salts when combined together.

The universal solvent, **water**, has a dielectric constant of 80. Therefore, when sodium chloride is dissolved in water, the dielectric constant of water reduces the electrostatic force, allowing the ions to move freely in the solution. They are also well-separated due to hydration with water **molecules**.

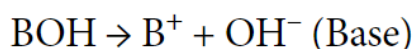
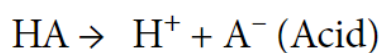


Ionization And Dissociation

Dissociation is the separation of ions from an ionic crystal when a solid ionic compound dissolves in water. On the other hand, ionization is the process where a neutral molecule breaks into charged ions when dissolved in a solution. The extent of ionization depends on the strength of the bonds between ions and the extent of solvation of ions.

Arrhenius Concept

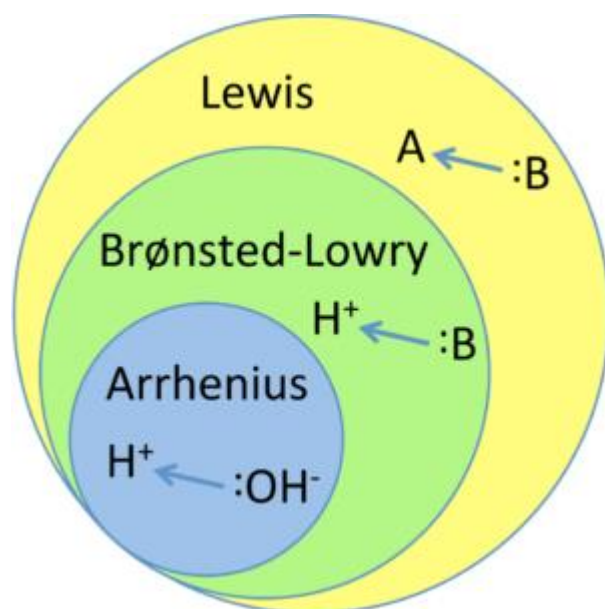
According to Arrhenius concept, Substances which produce H^+ ions when dissolved in water are called acids while those which ionize in water to produce OH^- ions are called bases.



Arrhenius proposed that acid-base reactions are characterized by acids if they dissociate in aqueous solution to form hydrogen ions (H^+) and bases if they form hydroxide (OH^-) ions in aqueous solution.

Bronsted-Lowry Concept

Bronsted and Lowry in 1923 independently proposed a more general definition of acids and bases. According to them, an acid is defined as any hydrogen-containing material (molecule, anion or cation) which can donate a proton to other substance and a Base is any substance (molecule, cation or anion) that can accept a proton from any other substance. Therefore, acids are proton donor whereas bases are proton acceptor.



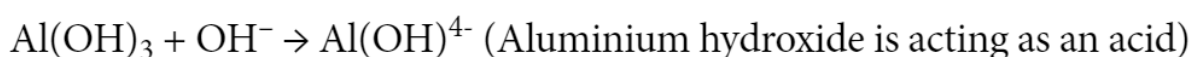
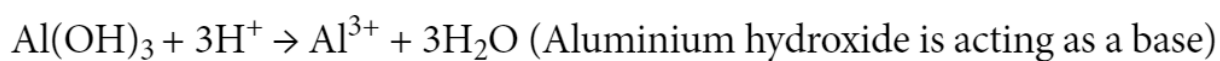
Acid-base pairs in which the members of reaction can be formed from each other by the gain or loss of protons are called conjugate acid-base pairs.

Lewis Concept

According to Lewis theory of acid-base reactions, bases donate pairs of electrons and acids accept pairs of electrons. Thus, it can be said that a Lewis acid is electron-pair acceptor.

The advantage of the Lewis theory is that complements the model of oxidation-reduction reactions. Oxidation-reduction reactions take place on a transfer of electrons from one atom to another, with a net change in the oxidation number of one or more **atoms**.

The Lewis theory further suggested that acids react with bases and share a pair of electrons but there is no change in the oxidation numbers of any atoms. Either an electron is transferred from one atom to another, or the atoms come together to share a pair of electrons.



These reactions are showing clearly: When Aluminium hydroxide accepts protons, it acts as a base. When it accepts electrons, it acts as an acid. This Lewis acid-base theory also explains why non-metal oxides such as **carbon** dioxide dissolve in H_2O to form acids, such as carbonic acid H_2CO_3 .

Solved Example for You

Question: Whether the following ions or molecules can act as Lewis acid or a Lewis base?

- Ag^+
- NH_3

Solution:

- A silver cation is Lewis acid
- Ammonia is Lewis base



**Carbon, Organic compounds
classification and nomenclature
of organic compound, isomerism
organic compound**

Introduction:

Carbon plays very important roles for all living beings.

The amount of carbon in the earth's crust is merely 0.02%, which is available in the form of minerals such as carbonates, hydrogen-carbonates, coal, and petroleum.

The presence of carbon in the atmosphere of the earth is 0.03%, in the form of carbon dioxide.

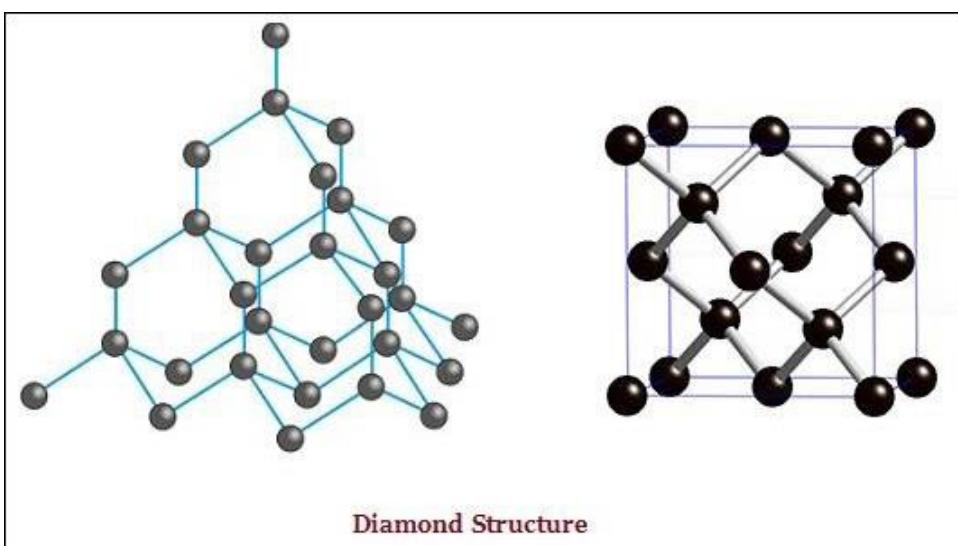
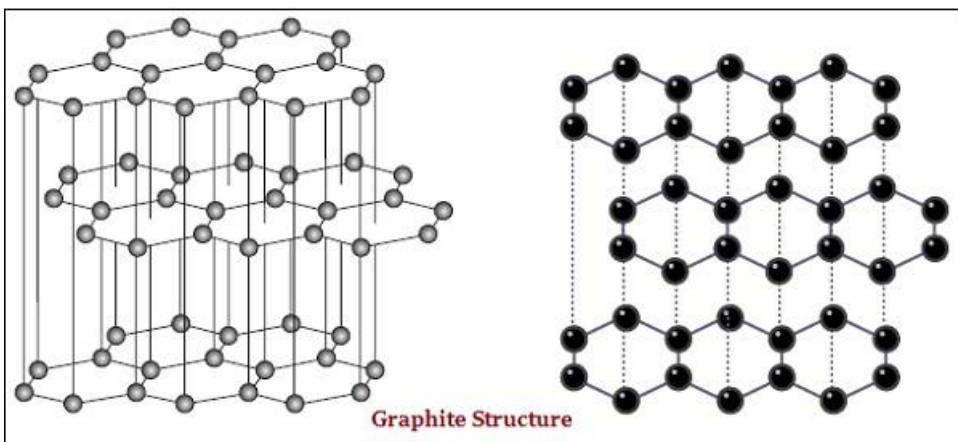
Compounds of Carbon

Almost all carbon compounds (except a few) are poor conductors of the electricity.

The diamond and graphite both are formed by carbon atoms; however, the difference lies between them in the manner in which the carbon atoms are bonded to one another.

In graphite, each atom of the carbon, is bonded to three other carbon atoms in the same plane, which gives a hexagonal array

In diamond, each atom of the carbon, is bonded to four other carbon atoms and form a rigid three-dimensional structure.

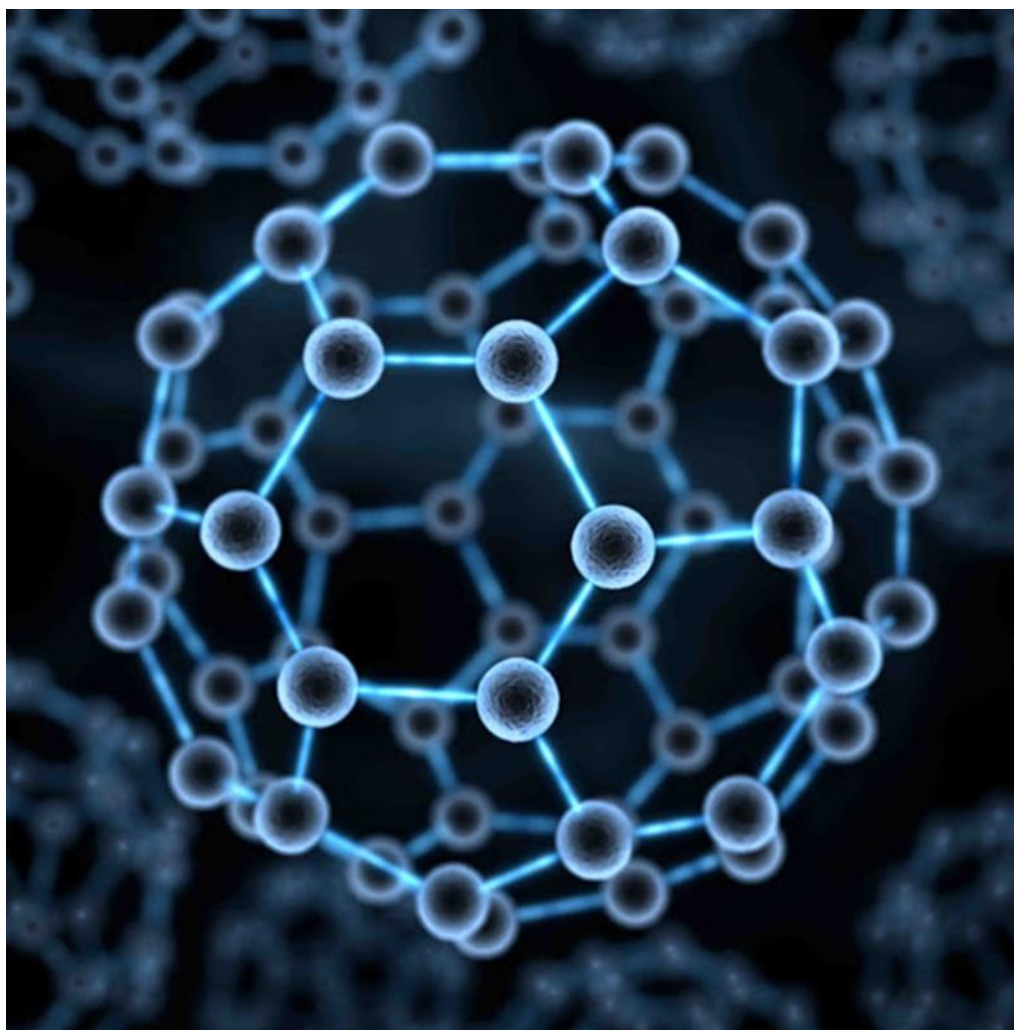


Allotropes of Carbon

Allotropy: The phenomenon in which the element exists in two or more different physical states with similar chemical properties are called Allotropy.

Carbon has Three Main Allotropes

- **Diamond:** In this, carbon, an atom is bonded to four other atoms of carbon forming three-dimensional structures. It is the hardest substance and an insulator. It is used for drilling rocks and cutting. It is also used for making jewellery.
- **Graphite:** In this, each carbon atom is bonded to three other carbon atoms. It is a good conductor of electricity and used as a lubricant.
- **Buckminster Fullerene:** It is an allotrope of the carbon-containing cluster of 60 carbon atoms joined together to form spherical molecules. It is dark solid at room temperature.



Physical Properties of Organic Compounds

Most of the organic compounds have low boiling and melting point, due to the weak force of attraction (i.e., the inter-molecular force of attraction) between these molecules.

Most carbon compounds are poor conductors of electricity, due to the absence of free electrons and free ions.

Compounds	M.P. (K)	B.P. (K)
Acetic acid (CH ₃ COOH)	290	391
Chloroform (CHCl ₃)	209	334
Ethanol (CH ₃ CH ₂ OH)	156	351
Methane (CH ₄)	90	111

Versatile Nature of Carbon: The existence of such a large number of organic compounds is due to the following nature of carbon,

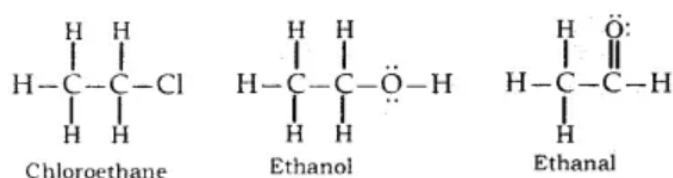
- Catenation
- Tetravalent nature.

(i) Catenation: The self linking property of an element mainly carbon atom through covalent bonds to form long straight, branched and rings of different sizes are called Catenation.

This property is due to

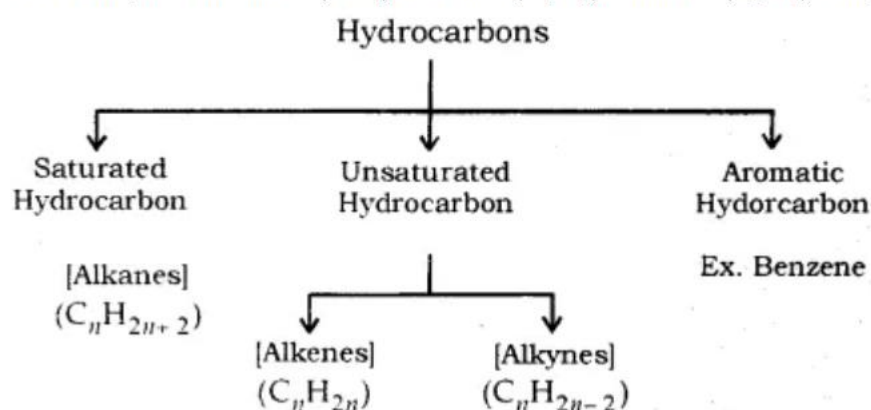
- The small size of the carbon atom.
- The great strength of the carbon-carbon bond.

(ii) Tetravalent Nature: Carbon has valency of four. It is capable of bonding with four other atoms of carbon or some other heteroatoms with single covalent bond as well as double or triple bond.

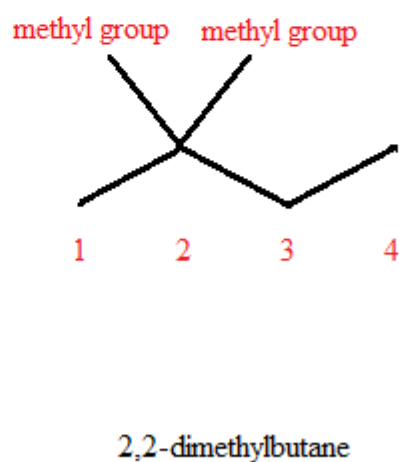
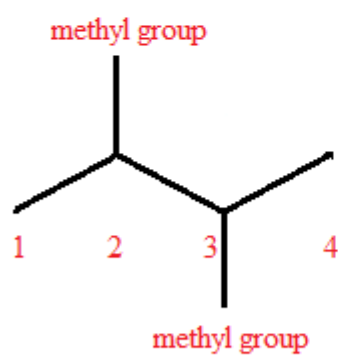
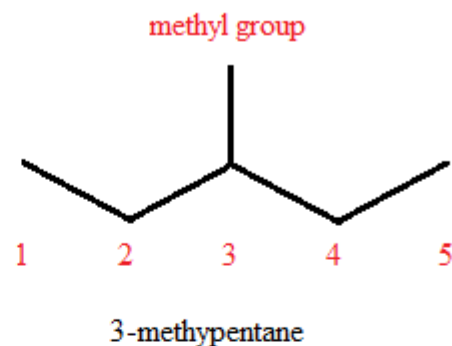
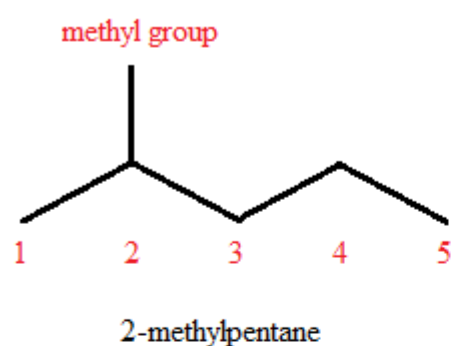



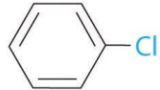
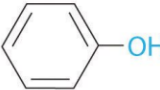




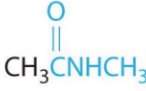
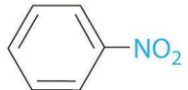
Hydrocarbons: Compounds of carbon and hydrogen are known as hydrocarbons.

For example; Methane (CH₄), Ethane (C₂H₆), Ethene (C₂H₄), Ethyne (C₂H₂) etc.



Number of 'C' atoms	Word root	IUPAC name	Structure	Molecular formula
1	Meth	Methane	CH ₄	CH ₄
2	Eth	Ethane	CH ₃ —CH ₃	C ₂ H ₆
3	Prop	Propane	CH ₃ —CH ₂ —CH ₃	C ₃ H ₈
4	But	Butane	CH ₃ —(CH ₂) ₂ —CH ₃	C ₄ H ₁₀
5	Pent	Pentane	CH ₃ —(CH ₂) ₃ —CH ₃	C ₅ H ₁₂
6	Hex	Hexane	CH ₃ —(CH ₂) ₄ —CH ₃	C ₆ H ₁₄
7	Hept	Heptane	CH ₃ —(CH ₂) ₅ —CH ₃	C ₇ H ₁₆
8	Oct	Octane	CH ₃ —(CH ₂) ₆ —CH ₃	C ₈ H ₁₈
9	Non	Nonane	CH ₃ —(CH ₂) ₇ —CH ₃	C ₉ H ₂₀
10	Dec	Decane	CH ₃ —(CH ₂) ₈ —CH ₃	C ₁₀ H ₂₂



Class	General Formula	Example	Common Name (Systematic Name)	Common Suffix/Prefix (Systematic)
Hydrocarbons				
Alkanes	RH	CH ₃ CH ₃	ethane	-ane
Alkenes	RR'C=CR''R'''	H ₂ C=CH ₂	ethylene (ethene)	-ene
Alkynes	RC≡CR'	HC≡CH	acetylene (ethyne)	(-yne)
Arenes	ArH ^a		benzene	-ene
Halogen-Containing Compounds				
Alkyl halides	RX	CH ₃ CH ₂ Cl	ethyl chloride (chloroethane)	halide (halo-)
Aryl halides	ArX ^a		chlorobenzene	halo-
Oxygen-Containing Compounds				
Alcohols	ROH ^a	CH ₃ CH ₂ OH	ethyl alcohol (ethanol)	-ol
Phenols	ArOH ^b		phenol	-ol
Ethers	ROR'	H ₃ CH ₂ COCH ₂ CH ₃	diethyl ether	ether
Aldehydes	RCHO		acetaldehyde (ethanal)	-aldehyde (-al)
Ketones	RR'C=O		acetone (2-propanone)	-one
Carboxylic acids	RCO ₂ H		acetic acid (ethanoic acid)	-ic acid (-oic acid)
Carboxylic Acid Derivatives				
Esters	RCO ₂ R'		methyl acetate (methyl ethanoate)	-ate (-oate)
Amides	RCONHR'		N-methylacetamide	-amide
Nitrogen-Containing Compounds				
Amines	RNH ₂ , RNHR', RNR'R''	CH ₃ CH ₂ NH ₂	ethylamine	-amine
Nitriles	RC≡N	H ₃ CC≡N	acetonitrile	-nitrile
Nitro compounds	ArNO ₂ ^a		nitrobenzene	nitro-

^aR indicates an alkyl group ^bAr indicates an *aryl* group.

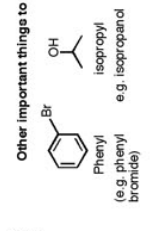
Homologous Series	Prefix or Suffix	Functional Group	Example
alkanes	--ane	$\begin{array}{c} \diagup \\ \text{C} - \text{H} \\ \diagdown \end{array}$	ethane C_2H_6
alkenes	--ene	$\begin{array}{c} \diagdown \quad \diagup \\ \text{C} = \text{C} \\ \diagup \quad \diagdown \end{array}$	ethene C_2H_4
haloalkanes	halo--	-Cl -Br -I	chloroethane $\text{CH}_3\text{CH}_2\text{Cl}$
alcohols	--ol hydroxyl--	- OH	ethanol or hydroxyethane $\text{CH}_3\text{CH}_2\text{OH}$
ethers	alkoxy--	- OR	methoxymethane CH_3OCH_3
aldehydes	--al	$\begin{array}{c} \text{O} \\ \parallel \\ -\text{C} \\ \diagdown \\ \text{H} \end{array}$	ethanal CH_3CHO
ketones	--one	$\begin{array}{c} \diagdown \quad \diagup \\ \text{C} = \text{O} \end{array}$	propanone CH_3COCH_3
carboxylic acid	--oic acid	$\begin{array}{c} \text{O} \\ \parallel \\ -\text{C} \\ \diagdown \\ \text{H} \end{array}$	ethanoic acid CH_3COOH
amines	amino-- --amine	- NH	aminomethane methylamine CH_3NH_2
amides	--amide	$\begin{array}{c} \text{O} \\ \parallel \\ -\text{C} \\ \diagdown \\ \text{NH} \end{array}$	ethanamide CH_3CONH_2
nitrils	--nitrile	- $\text{C}\equiv\text{N}$	propanenitrile $\text{CH}_3\text{CH}_2\text{CN}$

Summary Sheet - Introduction to Chemical Reactivity, Nomenclature, Boiling Points, and Water Solubility

Reactant #1	Reactant #2	Product	Example	Notes	Name	Type of Interaction	Found in	Example	Strength	Notes
Alkene	Pd/C + H ₂	Alkane		cis addition (hydrogens go on same side of alkene)	Ionic	Attraction between oppositely charged H and negative charged O, N or F.	Salts	NH ₄ ⁺ Cl ⁻	Strongest	Boiling point increases with strength of the intermolecular interactions. Water solubility increases with polarity.
Amine	Acid	Ammonium salt		amines but NOT amides (amides are not basic on nitrogen)	Hydrogen Bonding	Attraction between positively charged H and negatively charged O, N or F.	Water, alcohols, carboxylic acids, amides, amines		2nd strongest	2nd greatest for effect on water solubility and boiling points
Carboxylic acid	Base	Salt (Carboxylate salt)			Dipole-Dipole	Attraction between dipole moments caused by differences in electronegativity	ketones, aldehydes, esters, alcohols, etc. - any molecule with a strongly electronegative element (O, N, F, Cl, Br)		2nd weakest	Increases as electronegativity difference increases 3rd greatest for effect on water solubility and boiling points
Carboxylic acid	Alcohol, acid, heat	Ester		We use the alcohol as solvent.	Van Der Waals (London forces)	Attraction between temporary dipoles	Hydrocarbons		Weakest	Increases with surface area (increasing length of carbon chain) worst for water solubility (least polar) best for solubility in non-polar solvents (e.g. pentane)
Ester	Water, acid, heat	Carboxylic acid		This is the reverse of the above reaction. Here we use water as solvent.						
Ester	Water, base	Carboxylic acid		This is called ester hydrolysis or saponification						
Alkene	Water, acid	Alcohol		alcohol forms on most substituted carbon (Markovnikov rule)						
Alkene	Strong acid	Alkyl halide		halide adds to most substituted carbon (Markovnikov rule) proceeds through carbocation						
Alkene	Br ₂	Dibromide		Results in trans product						
Alcohol	K ₂ Cr ₂ O ₇	Carboxylic acid (primary alcohol) OR ketone (secondary alcohol)		Note that secondary alcohols stop at the ketone stage						
Alcohol	KMnO ₄	Carboxylic acid (primary alcohol)		Free-radical reaction (number of new C-C bonds depends on # of equivalents)						
Alkene	Cl ₂ , hv (or peroxides)	Alkyl chloride		Also gives 1,4 (para) product but never 1,3 (meta) product						
Benzene derivative	Br ₂ , FeBr ₃	Aryl bromide								
Thiol	"Oxidant"	Disulfide								

# Carbons	Root
1	Meth-
2	Eth-
3	Prop-
4	But-
5	Pent-
6	Hex-
7	Hept-
8	Oct-
9	Non-
10	Dec-

Primary: attached to ONE carbon atom
 Secondary: attached to TWO carbon atoms
 Tertiary: attached to THREE carbon atoms
 Quaternary: attached to FOUR carbon atoms

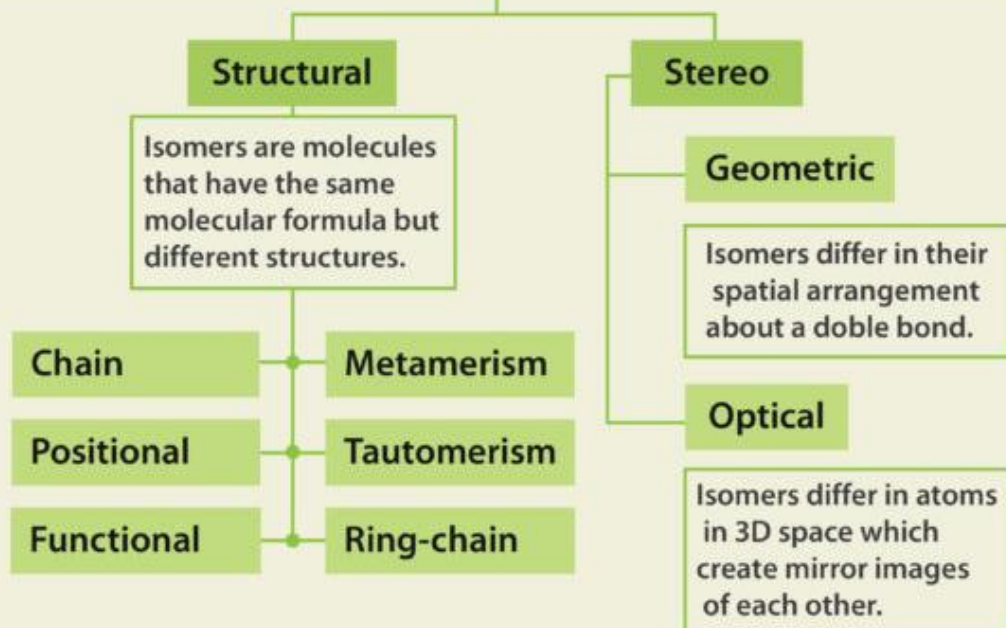


Functional Group	Name	Example	Name
R-	Alkyl		Pentane
-OH	Hydroxyl		Pentanol or pentyl alcohol
-Cl, -Br, -F, -I	Halide		Pentyl chloride
-NH ₂	Amine		Pentylamine
R'-O-R	Ether		Pentyl methyl ether
-SH	Thiol		Pentane thiol
	Alkene		Pentene
	Aldehyde		Pentanal
	Ketone		Butyl methyl ketone OR 2-pentanone
	Carboxylic acid		Pentanoic acid
	Ester		Methyl pentanoate
	Amide		N-methyl pentanamide

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 James A. Kleinburg
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 james@write4em.com

Isomerism

Isomers are molecules that have the same molecular formula, but have a different structures or arrangement of the atoms in space. This phenomenon is called Isomerism.



Electronic factors that influence organic reactions

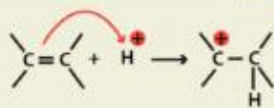
Inductive effect

Permanent state of polarization.



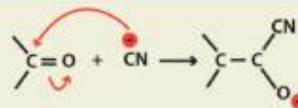
Electromeric effect

temporary polarisation and reversible, Occurs in the presence of reagent.



+E effect

π-electrons of the multiple bond are transferred to that atom to which the reagent gets attached



-E effect

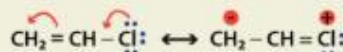
π-electrons of the multiple bond are transferred to that atom to which the attacking reagents do not get attached.

Mesomeric effect

The permanent polarization of a group conjugated with a π bond or a set of alternate π bonds.



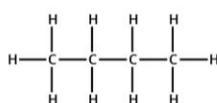
-M effect



+M effect

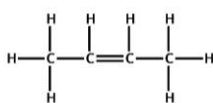
STRUCTURAL ISOMERISM

CHAIN



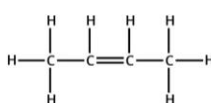
BUTANE

POSITION

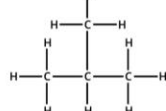


BUT-2-ENE

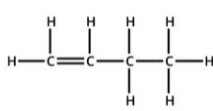
FUNCTIONAL



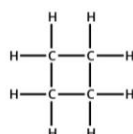
BUT-2-ENE



2-METHYL PROPANE



BUT-1-ENE



CYCLOBUTANE

DIFFERENT ARRANGEMENT OF A MOLECULE'S CARBON SKELETON

The positions of the carbon atoms in the molecule can be rearranged to give 'branched' carbon chains coming off the main chain. The name of the molecule changes to reflect this, but the molecular formula is still the same.

THE DIFFERING POSITION OF THE SAME FUNCTIONAL GROUP IN THE MOLECULE

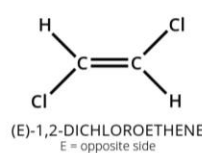
The molecular formula remains the same; the type of functional group also remains the same, but its position in the molecule changes. The name of the molecule changes to reflect the new position of the functional group.

DIFFERING POSITIONS OF ATOMS GIVE A DIFFERENT FUNCTIONAL GROUP

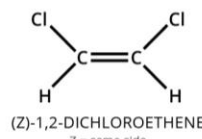
Also referred to as functional group isomerism, these isomers have the same molecular formula but the atoms are rearranged to give a different functional group. The name of the molecule changes to reflect the new functional group.

STEREOISOMERISM

GEOMETRIC



(E)-1,2-DICHLOROETHENE
E = opposite side

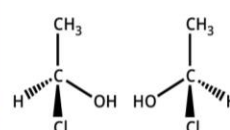


(Z)-1,2-DICHLOROETHENE
Z = same side

DIFFERENT SUBSTITUENTS AROUND A BOND WITH RESTRICTED ROTATION

Commonly exhibited by alkenes, the presence of two different substituents on both carbon atoms at either end of the double bond can give rise to two different, non-superimposable isomers due to the restricted rotation of the bond.

OPTICAL



L: (S)-1-CHLOROETHANOL
R: (R)-1-CHLOROETHANOL



NON-SUPERIMPOSABLE MIRROR IMAGES OF THE SAME MOLECULE

Optical isomers differ by the placement of different substituents, around one or more atoms in a molecule. Different arrangements of these substituents can be impossible to superimpose - these are optical isomers.