



Lecture One
Introduction to chemistry (matter , structure of atom , periodic table , isotopes , atomic number ,mass number, composition of matter ,types of bonds)

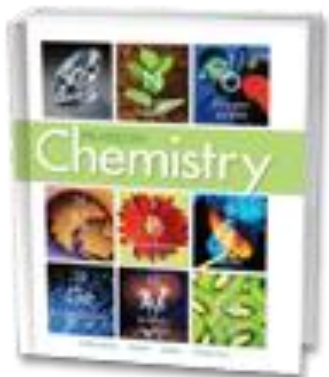
First Year

Medical Laboratory Techniques Department

Subject Lecturer

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What is Chemistry?

Chemistry : Is a whole branch of science about matter, which is anything that has mass and occupies space.

Chemistry is the study of the composition and properties of matter and the changes it undergoes.

Branches of Chemistry:



1- Organic Chemistry involves the study of the structure, properties, and preparation of chemical compounds that consist primarily of carbon and hydrogen.

2-Inorganic Chemistry is the study of the properties and behaviour of inorganic compounds. It covers all chemical compounds except organic compounds. Inorganic chemists study things such as crystal structures, metals, catalysts, and most elements in the Periodic Table.

3- Analytical Chemistry: involves the qualitative and quantitative determination of the chemical components of substances.

4- Physical Chemistry: The study of the effect of chemical structure on the physical properties of a substance.

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

MATTER

Matter: Is anything that occupies space and has weight. If you look around you, you will see matter. The table, books, walls, and your body are all composed of matter.

Physical States of Matter.

In general, we can group all matter into three groups called states of matter .

(1) Solids.

Solids have a definite shape and volume. Examples of solids are books, rocks, pieces of steel, and sand .

(1) Liquids.

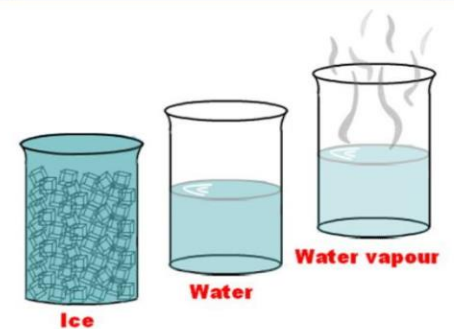
Liquids have a definite volume but indefinite shape. That is ,they take the shape of their container. Water, mercury, alcohol, and oils are liquids .

(1) Gases.

Gases have neither a definite shape nor a definite volume. They assume not only the shape of their container, but also the volume of their container.

Gases may be expanded or compressed to fit the container in which they are being placed. Therefore, the air in an automobile tire would, if released, expand to fill a large weather balloon.

States of Matter



Properties of Matter

Matter possesses two types of properties, physical and chemical. From the physical and chemical properties exhibited by a substance, it is possible to isolate, identify, and classify the particular substance.

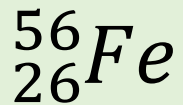
- ❑ Physical properties can be observed or measured without changing the composition of matter.
- ❑ Physical properties are used to observe and describe matter. Physical properties include: smell, color, shape, melting point, boiling point, density, solubility, and polarity.
- ❑ Chemical properties are any of the properties of matter that may only be observed and measured by performing chemical reaction. Examples of Chemical Properties
 - ✓ Reactivity with other chemicals
 - ✓ Toxicity
 - ✓ Coordination number
 - ✓ Flammability (is the ability of a substance to burn or ignite.
 - ✓ Heat of combustion (The heat of combustion is the total energy released as heat when a substance undergoes complete combustion with oxygen under standard conditions)

Energy

Energy : is the capacity of a physical system to perform work. However, it's important to keep in mind that just because energy exists, it doesn't mean it's necessarily available to do work. So, there are two types of energies: Kinetic Energy and Potential Energy.

- **Atom** : the smallest component of an element having the chemical properties of the element, consisting of a nucleus containing combinations of neutrons and protons and one or more electrons bound to the nucleus by electrical attraction.
- **Compound**: It is a substance made of more than one type of atoms. They are usually formed by a chemical process and atoms are bound together by chemical bonds.
- **Molecules** - This is the smallest unit of a compound. For example, water is dihydrogen oxide.
- **Atomic number (Z)**: The number of protons in the nucleus of an atom. In electrically neutral atoms, this number is also equal to the number of electrons orbiting about the atom's nucleus. For example: H = 1, Br = 35, Cl = 17
- **Mass numbers (A)**: is the total number of protons and neutrons in an atomic nucleus. For example

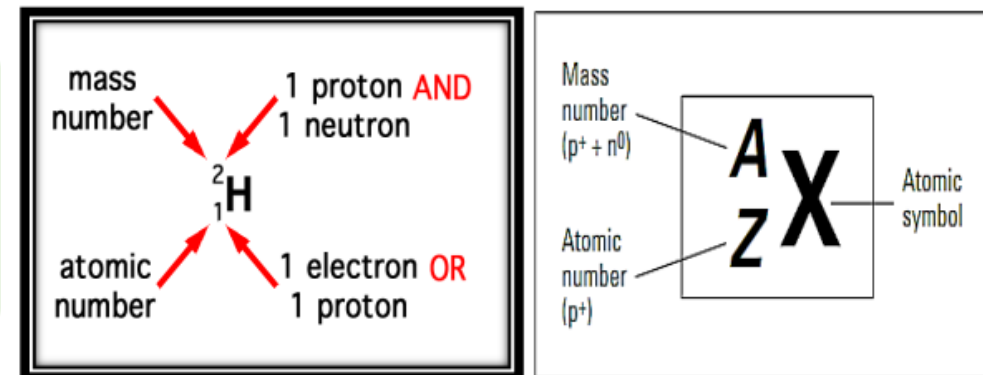
The iron nucleus which has 26 protons and 30 neutrons, is denoted as



Where the total atomic number is ($Z = 26$) and the mass number ($A = 56$).

The number of neutrons is simply the difference ($N = A - Z$)

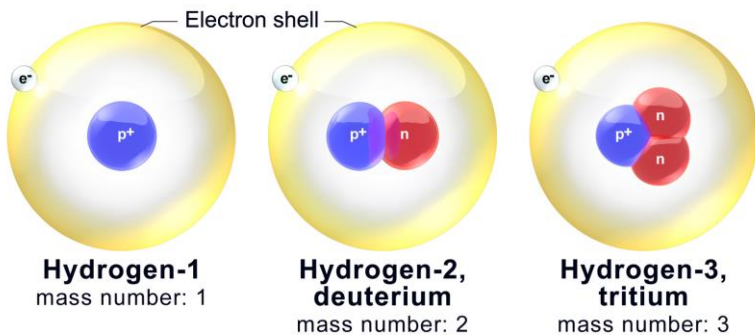
$$N = 56 - 26 = 30$$



Isotopes

Isotopes :- The atoms of a particular element can have an identical number of protons and electrons but varying numbers of neutrons.

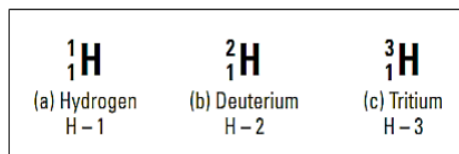
- If they have different numbers of neutrons, then the atoms are called isotopes (Atoms of the same element that have same atomic number but different numbers of neutrons).
- The figure below shows the three hydrogen isotopes, Hydrogen or (Protium) which has (1p + 0N), Deuterium which has (1p + 1N) and Tritium which has (1p + 2N).



Each of these isotopes have different abundance in nature. The table below outlines the average isotopic abundances of elements that are most commonly measured for stable isotope measurements

Other examples

Element	Isotopes
Carbon (C)	$^{12}_6\text{C}$, $^{13}_6\text{C}$, $^{14}_6\text{C}$
Nitrogen (N)	$^{14}_7\text{N}$, $^{15}_7\text{N}$
Oxygen (O)	$^{16}_8\text{O}$, $^{17}_8\text{O}$, $^{18}_8\text{O}$
Chlorine (Cl)	$^{35}_{17}\text{Cl}$, $^{37}_{17}\text{Cl}$
Uranium (U)	$^{235}_{92}\text{U}$, $^{238}_{92}\text{U}$, $^{239}_{92}\text{U}$
Sulphur (S)	$^{32}_{16}\text{S}$, $^{33}_{16}\text{S}$, $^{34}_{16}\text{S}$, $^{36}_{16}\text{S}$



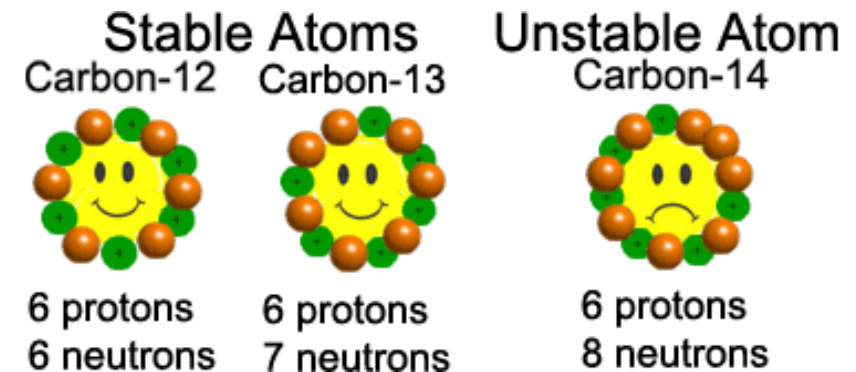
Hydrogen Isotopes

Hydrogen	Carbon	Nitrogen	Oxygen
^1H – 99.984%	^{12}C – 98.89%	^{14}N – 99.64%	^{16}O – 99.763%
^2D – 0.0156%	^{13}C – 1.11%	^{15}N – 0.36%	^{17}O – 0.0375%
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Stable vs. Unstable Isotopes

- In chemistry an isotope is a different form of an element which will have different neutron number and thus has a different mass number as well. But the chemical properties and other related characteristics remain same as that of the element.
- There are two types of isotopes, one which is found in nature, stable and not radioactive, the second type is the one which is naturally radioactive and not stable. Any element having excess of neutrons in the nucleus as compared to the stable form of atom tends to be unstable in nature.
- These elements tend to be radioactive and undergo radioactive decay by emitting energy

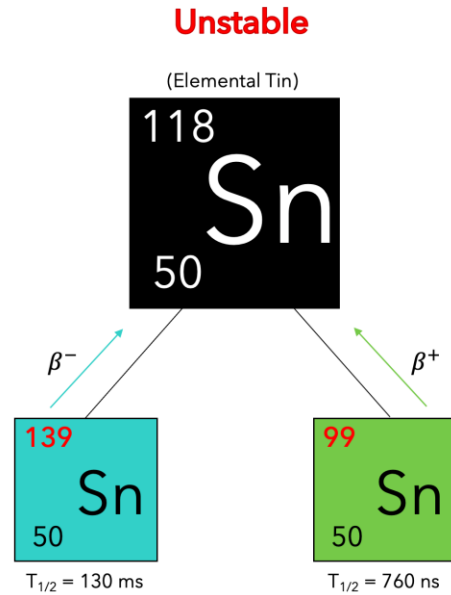
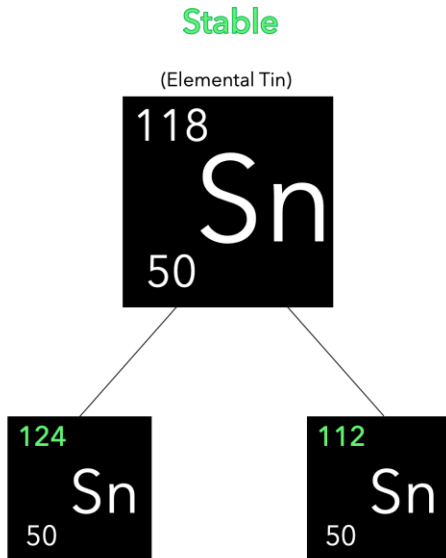
*Unstable nuclei decay by giving off Alpha particles (helium nuclei), Beta particles (electrons or antielectrons) or Gamma particles (photons of high energy electromagnetic radiation).



The examples of naturally occurring stable and radioactive or unstable isotopes are mentioned below:

There are two kinds of isotopes: *Stable* or *Unstable*

Tin has **10** stable isotopes.



Naturally occurring stable isotopes	Radioactive unstable isotopes
Carbon 12	Carbon 14
Calcium 40 Calcium 43 Calcium 46	Calcium 48
Vanadium 51	Vanadium 50
Germanium 70 Germanium 72 Germanium 74	Germanium 76
Selenium 74 Selenium 76 Selenium 78 Selenium 80	Selenium 82

Average Atomic Mass

The calculation of the average atomic mass of an atom is performed using the relative abundance data from the isotope of each atom. The average atomic mass of an element is the sum of the masses of its isotopes, each multiplied by its natural abundance (the decimal associated with percent of atoms of that element that are of a given isotope).

$$[(\% \text{ abundance of isotope}) (\text{mass of isotope})] + [(\% \text{ abundance of isotope}) (\text{mass of isotope})] + \dots / 100$$

For example, the natural abundance for chloro isotopes is 75.78% ^{35}Cl (34.968853 amu*) and 24.22% while ^{37}Cl (36.9659033 amu*). Calculate the atomic mass of Cl

$$\text{Average atomic mass} = [(75.78)(34.968853)] + [(24.22)(36.9659033)] / 100 = 35.45$$

Example 2 The natural abundance for boron isotopes is 19.9% ^{10}B (10.013 amu*) and 80.1% ^{11}B (11.009 amu*). Calculate the atomic mass of boron.

Ans. Average atomic mass = $[(19.9\%)(10.013)] + [(80.1\%)(11.009)] / 100 = 10.811$
(note that this is the value of atomic mass given on the periodic table)

Note: amu is the atomic mass unit, which is defined as 1/12th the mass of a carbon-12 atom. This value provides a reference point for measuring relative atomic masses.
The average atomic mass also known as the molar mass or relative atomic mass (RAM).

Tables of Isotopic Masses and Natural Abundances These tables list the mass and percent natural abundance for the stable nuclides.

Element	Symbol	Nominal Mass	Exact Mass	Abundance %
Hydrogen	H	1	1.00783	99.99
	D or ^2H	2	2.0141	.01
Carbon	C	12	12	98.91
		13	13.0034	1.09
Nitrogen	N	14	14.0031	99.6
		15	15.0001	0.37
Oxygen	O	16	15.9949	99.76
		17	16.9991	0.037
		18	17.9992	0.2
Fluorine	F	19	18.9984	100
Silicon	Si	28	27.9769	92.28
		29	28.9765	4.7
		30	29.9738	3.02
		31	30.9738	100
Phosphorous	P	31	30.9738	100
Sulphur	S	32	31.9721	95.02
		33	32.9715	0.74
		34	33.9679	4.22
		35	34.9689	75.77
Chlorine	Cl	37	36.9659	24.23
		35	34.9689	75.77
Bromine	Br	79	78.9183	50.5
		81	80.9163	49.5
Iodine	I	127	126.9045	100

What Is Chemical Bonding?

Chemical bonding refers to the formation of a chemical bond between two or more atoms, molecules or ions to give rise to a chemical compound. These chemical bonds are what keep the atoms together in the resulting compound.

The attractive force which holds various constituents (atoms, ions, etc.) together and stabilises them by the overall loss of energy is known as chemical bonding.

The main types of chemical bonds are ionic bond, covalent bond, hydrogen bond, and metallic bond



1. Ionic Bond

As the name suggests, ionic bonds are a result of the attraction between ions. Ions are formed when an atom loses or gains an electron. These types of bonds are commonly formed between a metal and a nonmetal

Examples

- Sodium (Na) and chlorine (Cl) combine to form stable crystals of sodium chloride (NaCl), also known as common salt.
- Magnesium (Mg) and oxygen (O) combine to form magnesium oxide (MgO).
- Potassium (K) and chlorine (Cl) combine to form potassium chloride (KCl).
- Calcium (Ca) and fluorine (F) combine to form calcium fluoride (CaF₂).

2. Covalent Bond

- In the case of a covalent bond, an atom shares one or more pairs of electrons with another atom and forms a bond. This sharing of electrons happens because the atoms must satisfy the octet (noble gas configuration) rule while bonding. Such a type of bonding is common between two nonmetals.
- The covalent bond is the strongest and most common form of chemical bond in living organisms. Together with the ionic bond, they form the two most important chemical bonds.
- A covalent bond can be divided into a nonpolar covalent bond and a polar covalent bond. In the case of a nonpolar covalent bond, the electrons are equally shared between the two atoms. On the contrary, in polar covalent bonds, the electrons are unequally distributed between the atoms.

Examples

- Two atoms of iodine (I) combine to form iodine (I₂) gas.
- One atom of carbon (C) combines with two atoms of oxygen (O) to form a double covalent bond in carbon dioxide (CO₂).
- Two atoms of hydrogen (H) combine with one atom of oxygen (O) to form a polar molecule of water (H₂O).

3. Hydrogen Bond

- A hydrogen bond is a chemical bond between a hydrogen atom and an electronegative atom. However, it is not an ionic or covalent bond but is a particular type of dipole-dipole attraction between molecules.
- First, the hydrogen atom is covalently bonded to a very electronegative atom resulting in a positive charge, which is then attracted towards an electronegative atom resulting in a hydrogen bond

Examples

- Hydrogen atom from one molecule of water bonds with the oxygen atom from another molecule. This bonding is quite significant in ice.
- In chloroform (CH_3Cl) and ammonia (NH_3), hydrogen bonding occurs between the hydrogen of one molecule and carbon/nitrogen of another.
- Nitrogen bases present in DNA are held together by a hydrogen bond.

4. Metallic Bond

A metallic bond is a force that holds atoms together in a metallic substance. Such solid consists of tightly packed atoms, where the outermost electron shell of each metal atom overlaps with a large number of neighboring atoms. As a consequence, the valence electrons move freely from one atom to another. They are not associated with any specific pair of atoms. This behavior is called non-localization. Examples: Sodium metal, Aluminum foil, Copper wire.

Types of Chemical Bonds

1. Ionic Bond

Metal atom loses electron(s) to nonmetal atom



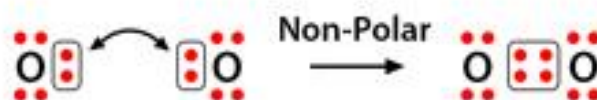
Sodium Chlorine Sodium chloride



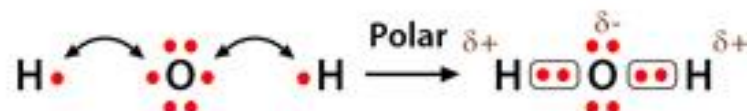
Magnesium Oxygen Magnesium oxide

2. Covalent Bond

Two nonmetal atoms share electrons



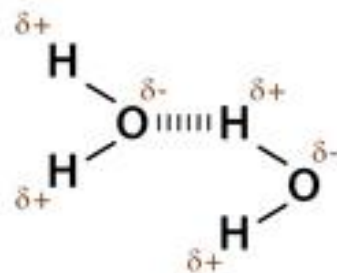
Oxygen atoms Oxygen molecule



Hydrogen Oxygen Hydrogen Water

3. Hydrogen Bond

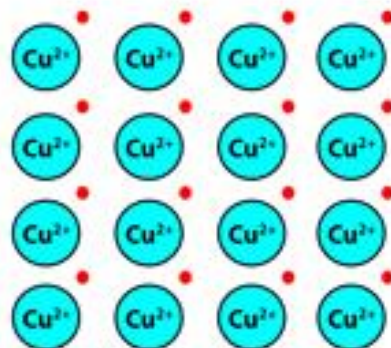
Hydrogen attracts an electronegative atom electrostatically



Two water molecules

4. Metallic Bond

Positive metal ions attract conducting electrons



Copper ions immersed in an electron cloud

The Periodic Table for PowerPoint

