Chapter 1 The structure of matter

Definition of matter

Anything which has mass and occupies space is called matter.

For example, book, pencil, water, air are composed of matter as we know that they have mass and they occupy space.

Classification of Matter

There are two ways of classifying the matter:

Physical classification and Chemical classification

(A) Physical Classification:

Matter can exist in three physical states: 1. Solids 2. Liquids 3. Gases

1. Solids: The particles are held very close to each other in an orderly fashion and there is not much freedom of movement. Solids have definite volume and definite shape.

2. Liquids: In liquids, the particles are close to each other but can move around. Liquids have definite volume but not definite shape.

3. Gases: In gases, the particles are far apart as compared to those present in solid or liquid states. Their movement is easy and fast.

Characteristics of Gases: Gases have neither definite volume nor definite shape. They completely occupy the container in which they are placed.

(B) Chemical Classification:

Based upon the composition, matter can be divided into two main types:

1. Pure Substances 2. Mixtures.

1. Pure substances: A pure substance may be defined as a single substance (or matter) which cannot be separated by simple physical methods.

Pure substances can be further classified as (i) Elements (ii) Compounds

(i) Elements: An element consists of only one type of particles. These particles may be atoms or molecules. For example, sodium, copper, silver, hydrogen, oxygen etc. are some examples of elements. They all contain atoms of one type. However, atoms of different elements are different in nature.

Some elements such as sodium . or copper contain single atoms held together as their constituent particles whereas in some others two or more atoms combine to give molecules of the element. Thus, hydrogen, nitrogen and oxygen gases consist of molecules in which two atoms combine to give the respective molecules of the element.

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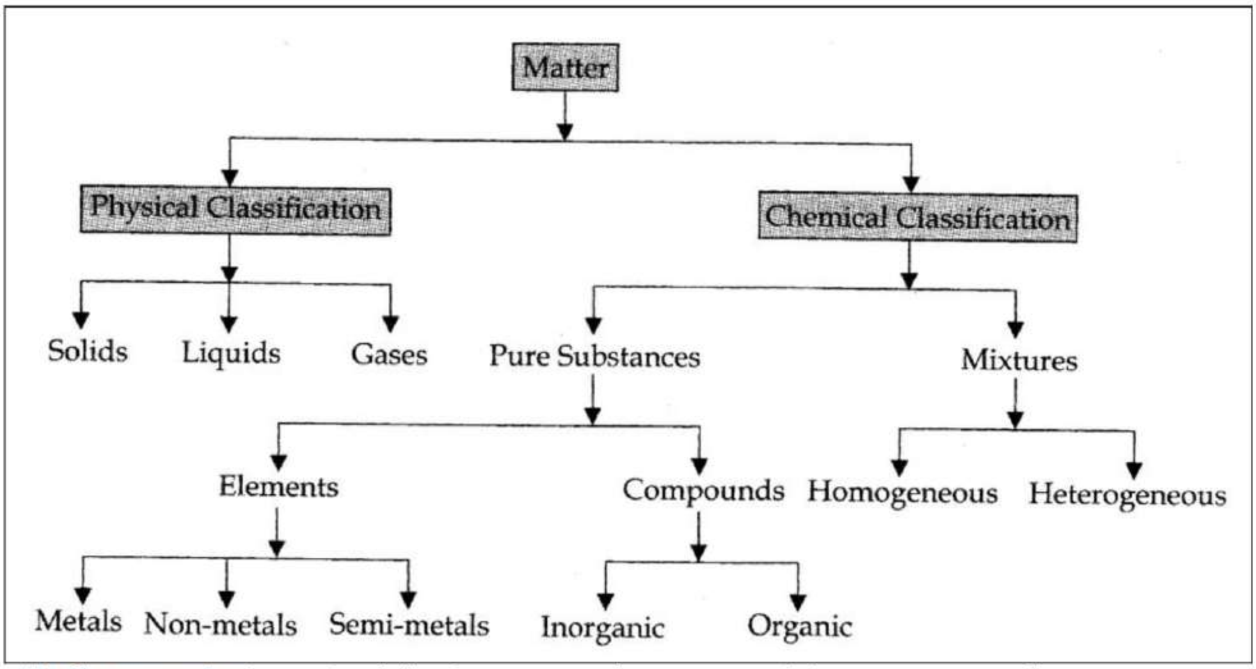

(ii) Compounds: It may be defined as a pure substance containing two or more elements combined together in a fixed proportion by weight and can be decomposed into these elements by suitable chemical methods. Moreover, the properties of a compound are altogether different from the constituting elements. The compounds have been classified into two types. These are:

(i) Inorganic Compounds: These are compounds which are obtained from non-living sources such as rocks and minerals. A few

examples are: Common salt, marble, gypsum, washing soda etc.

(ii) Organic Compounds are the compounds which are present in plants and animals. All the organic compounds have been found to contain carbon as their essential constituent. For example, carbohydrates, proteins, oils, fats etc.

2. Mixtures: The combination of two or more elements or compounds which are not chemically combined together and may also be present in any proportion, is called mixture. A few examples of mixtures are: milk, sea water, petrol, lime water, paint glass, cement, wood etc.

Types of mixtures: Mixtures are of two types:

(i) Homogeneous mixtures: A mixture is said to be homogeneous if it has a uniform composition throughout and there are no visible boundaries of separation between the constituents.

For example: A mixture of sugar solution in water has the same sugar water composition throughout and all portions have the same sweetness.

(ii) Heterogeneous mixtures: A mixture is said to be heterogeneous if it does not have uniform composition throughout and has visible boundaries of separation between the various constituents. The different constituents of a heterogeneous mixture can be seen even with naked eye.

Avogadro's number or Avogadro's constant (N_A)

One gram atom of any element contains the same number of atoms and one gram molecule of any substance contains the same number of molecules.

The value was found to be 6.022137×10^{23}

The value generally used is 6.022×10^{23}.

This is called Avogadro's number or Avogadro's constant (N_A)

Avogadro's number is a fundamental constant in chemistry that represents the number of entities (atoms, molecules, ions, etc.) in one mole of substance.

A mole of hydrogen atom means 6.022×10^{23} atoms of hydrogen whereas a mole of hydrogen molecule means 6.022×10^{23} molecules of hydrogen or $2 \times 6.022 \times 10^{23}$ atoms of hydrogen .

A mole of oxygen molecule means 6.022×10^{23} molecules of oxygen or $2 \times 6.022 \times 10^{23}$ atoms of oxygen.

Mole

A mole is defined as that amount of substance which contains Avogadro's number of atoms if the substance is atomic or Avogadro's number of molecules if the substance is molecular.

1 mole of carbon atoms = 6.022×10^{23} atoms of carbon.

1 mole of sodium atom = 6.022×10^{23} atoms of sodium

1 mole of Oxygen atom = 6.022×10^{23} atoms of oxygen

1 mole of Oxygen molecule = 6.022×10^{23} molecules of oxygen

1 mole of water = 6.022×10^{23} molecules of water

In case of gases, a mole is defined as that amount of the gas which has a volume of 22.4 litres at STP.

1 mole of Oxygen gas = 22.4 litres of oxygen at STP

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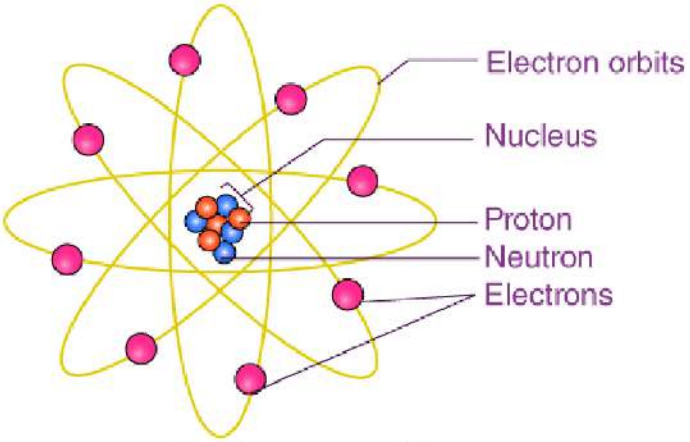
What Is Atomic Structure?

The atomic structure of an element refers to the constitution of its nucleus and the arrangement of the electrons around it. Primarily, the atomic structure of matter is made up of protons, electrons and neutrons.

The **protons and neutrons** make up the nucleus of the atom, which is surrounded by the electrons belonging to the atom. The **atomic number** of an element describes the total number of protons in its nucleus.



The Structure of Atom

Neutral atoms have equal numbers of protons and electrons. However, atoms may gain or lose electrons in order to increase their stability, and the resulting charged entity is called an ion.

Atoms of different elements have different atomic structures because they contain different numbers of protons and electrons. This is the reason for the unique characteristics of different elements.

Subatomic Particles

Protons

- Protons are positively charged subatomic particles. The charge of a proton is 1e, which corresponds to approximately 1.602×10^{-19}C
- The mass of a proton is approximately 1.672×10^{-24}g
- 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.

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Neutrons

- The mass of a neutron is almost the same as that of a proton, i.e., 1.674×10^{-24}g
- 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}C
- The mass of an electron is approximately 9.1×10^{-31} Kg.
- Due to the relatively negligible mass of electrons, they are ignored when calculating the mass of an atom.

In a neutral atom, the number of protons equals the number of electrons, although, the number of neutrons may differ according to the isotope. As a result, it becomes a necessity to provide a system of representation that can distinguish between the isotopes as well as different elements simultaneously.

Symbolic Representation of Atoms

An atomic symbol consists of three parts represented as: The mass number is the sum of the number of neutrons and protons. Atomic number is the number of protons. The electrical charge represents the net gain (for anions) or loss (for cations) of electrons.

Some of the important points that need to be remembered are:

1. The symbol may be written without the atomic number Z, as it is characteristic of each element and no two elements have the same alphabetical symbol X.
2. The symbol X can either contain a single alphabet, represented with CAPITAL LETTERS or two alphabets, represented with a CAPITAL LETTER followed by a "small letter".
3. The mass number must be included in the symbol to identify the isotope.
4. The representation of the element without mass number usually denotes the collection of the atoms of an element in the order of its natural abundance.

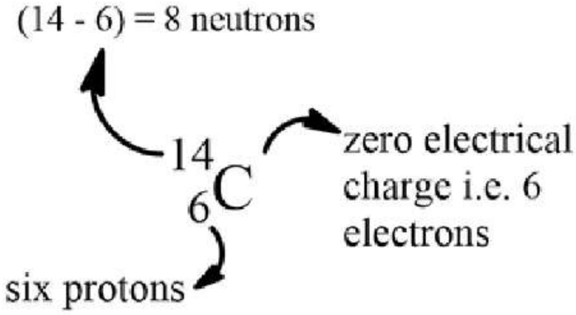
5. Electrical charge is not needed for representation of a neutral atom since there is neither excess nor lack of electrons in it. However, it is needed for an ion.

Utility Of The Atomic Symbols

1. The atomic symbols are useful in determining the number of protons, neutrons, and electrons.
2. The atomic symbols are useful in identifying the components of a compound.
3. The atomic symbols are useful in identifying the group and the period to which the element belongs.
4. The atomic symbols are useful in determining the electronic configuration.

The number of protons, neutrons, and electrons can be easily obtained by the following method.

Let’s look at the following:



The carbon-12 has 6 protons, six neutrons, and 6 electrons since the number of protons & electrons must be equal for it to be neutral.

While calculating the number of electrons in an ion, we should subtract the positive charge (for cation) from the number of protons and add the negative charge (for anion) to the number of protons.

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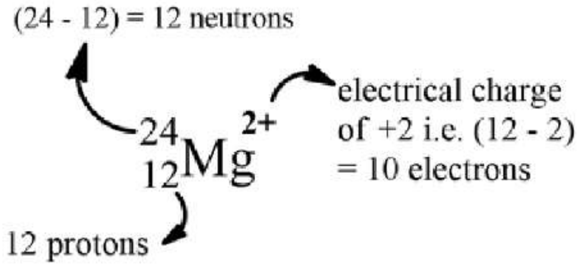
Magnesium ion (formed from Mg – 24)



The Atomic Mass Unit (U Or U.M.A.)

The reference mass is that of the 12C isotope of carbon, to which a molar mass of 12 g is assigned. This means that 12 g of this element contains a number of atoms equal to Avogadro's number ($N_A = 6{,}023 \times 10^{23}$). To express the molar masses of other atoms, a reference mass equal to 1/12 (which is 1 g) of that of carbon is used. Subsequently, the atomic mass unit (u.m.a.) is defined as the quotient of 1 g/N_A (where N_A is Avogadro's number). This unit is also known as the Dalton (Da), although this term is primarily employed to define the molar masses of macromolecules.

1 u (or 1 u.m.a. or 1 Dalton) = $1{,}6605 \times 10^{-27}$ kg

The Mass of Atom:

The mass of an atom is approximately equal to the sum of the masses of the particles it comprises: $m_{atome} \approx Z.m_p + (A-Z).m_n + Z.m_{e-}$

The isotopes

An isotope is one of two or more species of atoms of a chemical element with the same atomic number and position in the periodic table and nearly identical chemical behavior but with different atomic masses and physical properties. Every chemical element has one or more isotopes.

Example: lithium:

- Lithium-6 atom: Has 3 protons, 3 neutrons (6-3), and 3 electrons.
- Lithium-7 atom: Has 3 protons, 4 neutrons (7-3), and 3 electrons.
- For example, carbon-14, carbon-13, and carbon-12 are all isotopes of carbon. Carbon-14 contains a total of 8 neutrons, carbon-13 contains a total of 7 neutrons, and carbon-12 contains a total of 6 neutrons.

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The mass defect and nuclear Binding Energy

The difference between the calculated and experimentally measured masses is known as the mass defect of the atom. The conversion between mass and energy is most identifiably represented by the mass–energy equivalence equation as stated by Albert Einstein: $E = mc^2$, where E is energy, m is mass of the matter being converted, and c is the speed of light in a vacuum. Using this mass–energy equivalence equation, the nuclear binding energy of a nucleus may be calculated from its mass defect. A variety of units are commonly used for nuclear binding energies, including electronvolts (eV), with 1 eV equaling the amount of energy necessary to move the charge of an electron across an electric potential difference of 1 volt: 1.602×10^{-19} J.

As a simple example of the energy associated with the strong nuclear force, consider the helium nucleus composed of two protons and two neutrons. The total mass of these subatomic particles may be calculated as:

$$\underbrace{(2x1.0073amu)}_{\textbf{Protons}}+\underbrace{(2x1.0087amu)}_{\textbf{neutrons}}=4.031amu$$

However, mass spectrometric measurements reveal that the mass of an He nucleus is 4.0026 amu, less than the combined masses of its constituent subatomic particles. This difference between the calculated and experimentally measured masses is known as the mass defect of the nucleus. In the case of helium, the mass defect indicates a “loss” in mass of 4.031 amu – 4.0026 amu = 0.0284 amu. The loss in mass accompanying the formation of a nucleus from protons and neutrons is due to the conversion of that mass into energy that is evolved as the nucleus forms. The nuclear binding energy is the energy produced when the atoms’ nucleons are bound together; this is also the energy needed to break a nucleus into its constituent protons and neutrons. In comparison to chemical bond energies, nuclear binding energies are *vastly* greater. Consequently, the energy changes associated with nuclear reactions are vastly greater than are those for chemical reactions.

The conversion between mass and energy is most identifiably represented by the mass-energy equivalence equation as stated by Albert Einstein: $E=mc^2$

where E is energy, m is mass of the matter being converted, and c is the speed of light in a vacuum. This equation can be used to find the amount of energy that results when matter is converted into energy. Using this mass-energy equivalence equation, the nuclear binding energy of a nucleus may be calculated from its **mass defect**, as demonstrated in Example. A variety of units are commonly used for nuclear binding energies, including electron volts (eV), with 1 eV equaling the amount of energy necessary to the move the charge of an electron across an electric potential difference of 1 volt, making $1eV=1.602\times10^{-19}J$.

The relative stability of a nucleus is correlated with its binding energy per nucleon, the total binding energy for the nucleus divided by the number or nucleons in the nucleus.

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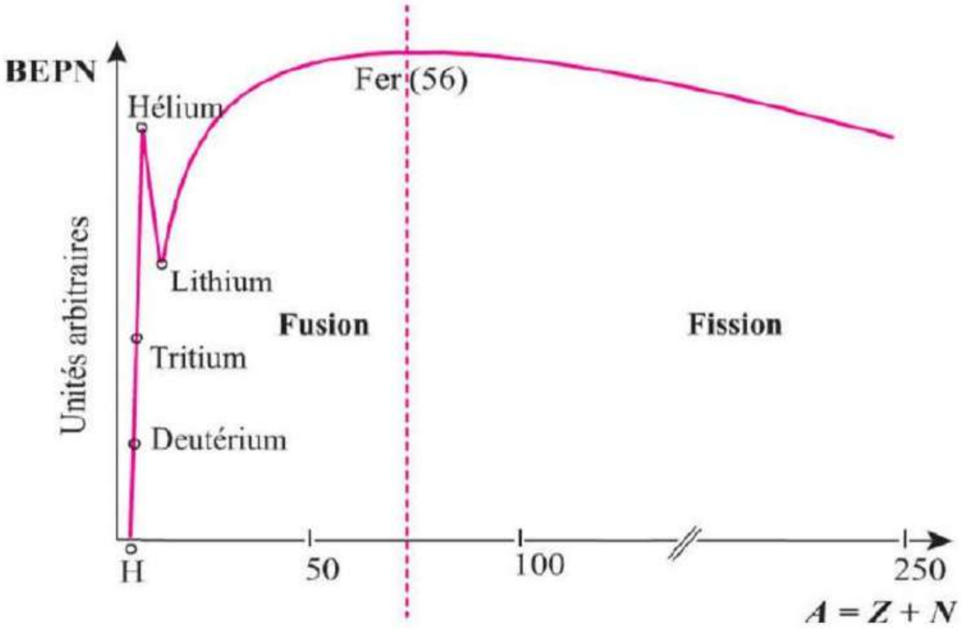
Aston curve



The Aston curve, in nuclear physics, is a graph representing the binding energy per nucleon of atomic nuclei as a function of their mass number.

Nuclear reactions have the effect of producing more stable elements, meaning those with higher binding energy per nucleon. There are two types of nuclear reactions. Fusion nuclear reactions occur in light elements such as hydrogen, generating heavier elements. They are currently difficult to control and occur, for example, in the explosion of hydrogen bombs. Nuclear fission reactions occur in heavy elements, generating lighter elements. They take place, notably, in nuclear power plants and atomic bombs.

The binding energy increases rapidly starting from hydrogen (no binding energy) and rises to the value of helium. The curve descends for lithium and then steadily increases, reaching its peak in the region of metals like iron and copper. It gradually descends as Z (the atomic number) increases. This allows for distinguishing the two major regions, separated by a dashed line in the figure.

Example: Calculation of Binding Energy per Nucleon

The iron nuclide ${}^{56}_{26}Fe$ lies near the top of the binding energy curve and is one of the most stable nuclides. What is the binding energy per nucleon (in MeV) for the nuclide ${}^{56}_{26}Fe$ (atomic mass of 55.9349 amu)?

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Solution

As in Example, we first determine the mass defect of the nuclide, which is the difference between the mass of 26 protons and 30 neutrons, and the observed mass of an Fe atom:

$$\text{Mass defect}=[(26\times 1.0073\,\text{amu})+(30\times 1.0087\,\text{amu})]-55.9349\,\text{amu}$$
$$=56.4508\,\text{amu}-55.9349\,\text{amu}=0.5159\ \text{amu}$$

We next calculate the binding energy for one nucleus from the mass defect using the mass-energy equivalence equation:

$$E=mc^2=0.5159\,\text{amu}\times(1.6605\times 10^{-27}\,\text{kg}/1\,\text{amu})\times(2.998\times 10^{8}\,\text{m/s})^2$$
$$=7.699\times 10^{-11}\,\text{kg}\cdot\text{m/s}^2=7.699\times 10^{-11}\,\text{J}$$

We then convert the binding energy in joules per nucleus into units of MeV per nuclide:

$$7.699\times 10^{-11}\,\text{J}\times(1\,\text{MeV}/1.602\times 10^{-13}\,\text{J})=480.58\,\text{MeV}$$

Finally, we determine the binding energy per nucleon by dividing the total nuclear binding energy by the number of nucleons in the atom:

$$\text{Binding energy per nucleon}=480.58\,\text{MeV}/56=8.581\,\text{MeV/nucleon}$$