Chapter 1 :

Matter, elements, atomes, isotopes, binding energy

What is chemistry ?

Chemistry is the study of composition, structure, properties and reactions of matter and the changes it undergoes. And the energy associated withe those chages.

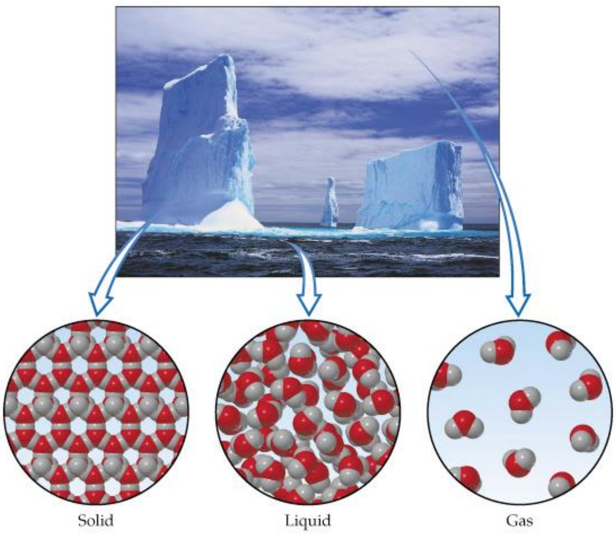
I. What is the matter ?

Matter is any thing that has mass and takes up space.

I.1. States of matter

Any matter can exist in one of 3 states : solid state, liquid state and gaseous state.

- In the solid state : the substances are rigid and have definite shapes, and the volumes do not vary much with changes in temperature and pressures.
- In the liquid state : the substances are not rigid and have no definite shapes, it takes the shape of the container in which it is placed in, and the volumes vary with changes in temperature and slightly compressed with pressures.
- In the gaseous state : the substances are not rigid and have no definite shapes and volume, it fills and takes the shape of the container in which it is placed in, and the volumes vary much with changes in temperature and compressed with pressures.



I.2. Classification of Matter:

matter can be classified according to its compositions: pure substances and mixtures.

I2.1. Pure Substance: a pure substance will always have the same composition and it cannot be more purified. Pure substances are either elements or compounds:

- **Element:** is a substance that consists of identical atoms (hydrogen, oxygen, and Iron). An element cannot be divided by chemical and physical methods. 116 elements are known (88 occur in nature and chemist have made the others in the lab).

Note: some elements consist of single atoms (for example, Ar and He). They are called **monatomic elements**. Some elements are **diatomic** and they consist of two atoms, connected to each other by a chemical bond (for example, N_2 and O_2). Some elements are **polyatomic** (many atoms) (for example, O_3 and S8).

- **Compound**: is a pure substance made up of two or more elements in a fixed ratio by mass (for example, water H_2O and Sodium chloride NaCl).

Formula: the formula gives us:

1- the ratios of the compound's constituent elements.

2- identifies each element by its atomic symbol.

Eg : H_2O, CO, CH_4

I.2.2. Mixture: is a combination of two or more pure substances. Mixtures are divided into two groups:

- **Homogeneous**: the mixture is uniform and throughout and no amount of magnification will reveal the presence of different substances (for example, air, and salt in water).
- **Heterogeneous**: the mixture is not uniform. Therefore, it contains regions that have different properties from those of other regions (for example, soup, milk, and blood).

I.3. Chemical and physical changes :

I.3.1. Chemical change (chemical reaction): when the substances are used up (disappear) and others are formed to take their place.

Example :

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Wood + fire $\longrightarrow$ Ashes

I.3.2. Physical change: when the state/appearance of the matter will change but its components will remain the same

I.4. Chemical and physical properties:

I.4.1. Chemical properties : are properties that matter exhibits as it undergoes changes in chemical composition.

I.4.2. Physical properties : are properties that can be observed in the absence of any chemical reaction. Color, density, hardness, melting point, boiling point, and electrical and thermal conductivitie are physical properties.

I.5. Extensive and Intensive Property

I.5.1. Intensive properties : are the same for all samples; do not depend on sample size; and include, for example, color, physical state, and melting and boiling points

I.5.2. Extensive properties : depend on the amount of material and include mass and volume. The ratio of two extensive properties, mass and volume, is an important intensive property called **density**.

I.6. Atomic Theory of Matter :

I.6.1. The *law of conservation of mass* : was discovered by the French chemist, Antoine Lavoisier. It states that atoms are neither created nor destroyed during a chemical change; the total mass of matter present when matter changes from one type to another remains constant.

I.6.2. the law of definite *proportions* or the *law of constant composition(* Joseph Proust)*:* all samples of a pure compound contain the same elements in the same proportion by mass. The suggestion that the numbers of atoms of the elements in a given compound always exist in the same ratio is consistent with these observations.

I.6.3. The *law of multiple proportions* (Dalton) : Dalton used data from Proust, as well as results from his own experiments, to formulate another interesting law. states that when two elements react to form more than one compound, a fixed mass of one element will react with masses of the other elements in a ratio of small, whole numbers.

II. Atoms:

Atoms contain up of 3 basic components known as subatomic particles, consisting of **protons** (positively charged), **neutrons** (no charge), and **electrons** (negatively charged). Protons and neutrons are in the center of the atom, and they make up the nucleus.

II.1. Proton Definition: Protons are positively charged subatomic particles, found in the nucleus of all atoms. The atomic number of an element is equal to the number of protons in the nucleus.

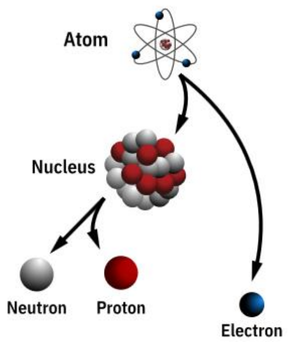
II.2. Neutron definition: A neutron is a undeniably a neutrally charged subatomic particle, found in the nucleus of all atoms except hydrogen.

II.3. Electron definition: Electrons are the subatomic particles that orbit the nucleus of an atom. They are negative in charge and are much smaller than protons or neutron. In fact, they are 1,800 times smaller. They also carry electricity.

Table 1 : Properties of Subatomic Particles

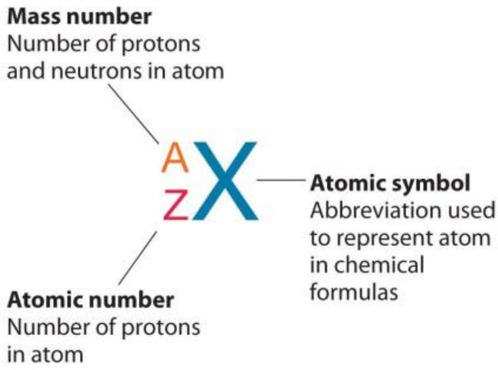
	Mass		Charge	
	Grams	Atomic Mass units (u)	Coulombs	Electronic charge
Proton	1.67×10^{-24}	1.007276	$+ 1.602 \times 10^{-19}$	1+
Neutron	1.67×10^{-24}	1.008665	0	0
Electron	9.11×10^{-28}	0.0005486	-1.602×10^{-19}	1-

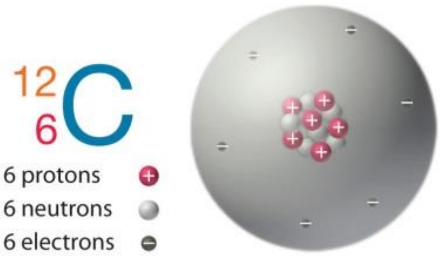
- **The masses of the proton and neutron are different in the fourth significant figure.**
- **The charges of the proton and electron, however, are believed to be exactly equal in magnitude (but opposite in sign).**
- **the mass of an atom is concentrated in its nucleus.**



where X represents the chemical element, A is the mass number, and Z is the atomic number.

For example, ${}^{12}_{6}C$ represents the carbon nucleus with six protons and six neutrons (or 12 nucleons).





- **Atomic number (Z)**: the number of protons in the nucleus of an atom.
- **Mass number (A)**: the number of protons (Z) + number of neutrons (N) in the nucleus of an atom.

$$A = Z + N$$

III. Avogadro numbr, mole concept and molar atomic :

III.1. Avogadro numbrand mole concept :

The mole, or "mol" is a unit of measurement in chemistry, used to designate a very large number of molecules, atoms, or particles. This very large number is called Avogadro's Number: 6.02214×10^{23}, the number of units in a mole.

Examples of Understanding Moles in Chemistry

1. Refer to the molecule $\mathbf{O_2}$

Since the coefficient of the compound is 1, we are working with 1 mole of the molecule O_2.

1 mole of $O_2 = 6.02214 \times 10^{23}$ molecules O_2

Since the atom O has a subscript 2, this means that there are two moles of O atoms in the molecule.

2 moles of O = $2 \times 6.02214 \times 10^{23}$ atoms = 12.04428×10^{23} oxygen atoms

2. Here's one more that's challenging: **3NaCl**

Since the coefficient of the compound is 3, we are working with 3 moles of the compound NaCl

3 moles of NaCl = 3 x 6.02214 x 10^{23} molecules NaCl

Within the compound there are 3 mol Na (the coefficient gets distributed to the atoms as well!)

3 moles of NaCl = 3 x 6.02214 x 10^{23} = 18.0663 x 10^{23} atoms of Na

Within the compound there are 3 mol Cl

3 moles of NaCl = 3 x 6.02214 x 10^{23} = 18.0663 x 10^{23} atoms of Cl

III.2. Molar mass :

- Mass of atom (from periodic table)

1 mole of atoms = gram atomic mass = 6.022×1023 atoms

Molecules

- Sum of atomic masses of all atoms in compound's formula

1 mole of molecule X = gram molecular mass of X

= 6.022 × 1023 molecules

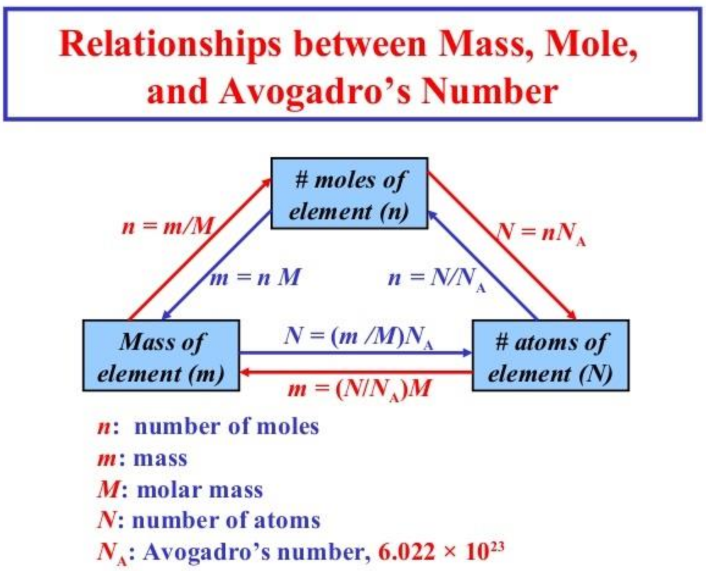
General

Molar mass (MM) is the mass of 1 mole of substance (element, molecule, or ionic compound) under consideration

1 mol of X = gram molar mass of X = 6.022 × 1023 formula units

- **Converting from Grams to Moles :**

It is possible to convert from grams to moles and vice versa using an element or compound's molecular weight. Recall the the molecular weight of a compound is the sum of the molecular weight of its elemental components.



Examples of Gram Conversions :

1. How many grams are in 9.2 moles of NO_2?

The molecular weight (mass) of NO_2 is 46 g/mol

To find the grams in 1.2 moles, divide by the molecular mass!

9.2/46 = 0.2 grams NO_2

2. How many atoms are in 1.5 grams of CO?

First convert the grams of CO to moles using molecular weight

The molecular mass of CO is 28 g/mol

Moles CO = 1.5 x 28 g/mol = 42 moles

To find the number of atoms, multiply the number of moles by Avogadro's number

Atoms CO = 42 moles x 6.02214 x 10^{23} = 252.93 x 10^{23} atoms CO

III.2. Atomic mass unit (amu):

Because atoms are too small to be seen and too small to have their masses measured individually on laboratory balance in a unit of grams, an atomic mass scale was devised in which mass is measured in atomic mass units (amu) (the SI symbol is u).

The atomic mass unit is defined as 1/12th of the mass of atom of carbon-12 isotope.

1 mole → N atoms of C-12 → 12g

1 atom of C-12 → X

X = 1 x 12 /N

And : 1uma = 1/12 (the masse of an atom of C-12) = 1/12 (12/ N)

So a unit of the scale relative masses of atoms (1 amu = 1.66×10-24 g).

For example :

1 mole of Na → N atoms of sodium → 23 g/mol

1 atom of sodium → X

X=23/6.023 10^{23}

X =

1 mole of Nacl → N moleculs of NaCl → g/mol

1 atom of NaCl → X

X=/6.023 10^{23}

X =

Density 1st lecture

Density is defined as the mass per unit volume.

density = mass/volume S.I. units for density = kg/m3

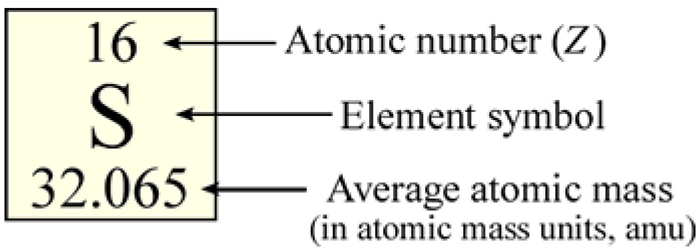
IV. sotopes :

IV.1. Isotopes: Atoms with same number of protons and electrons but different numbers of

neutrons. For example carbon: carbon-12, carbon-13 and carbon-14 (which have 6, 7, 8 neutrons but all of them have 6 protons and 12 electrons). The isotopes of the same element have almost the same properties.

IV.2. Isotopic Masses, Percent Natural Abundance, and Weighted-Average Atomic Mass :

Because most elements occur as isotopes and different isotopes have different masses, the atomic mass of an element is the average of the isotopic masses, weighted according to their naturally occurring abundances; this is the mass of each element recorded on the periodic table, also known as the relative atomic mass (A_r).



Treating isotopic masses in weighted averages gives greater importance to the isotope with greatest percent natural abundance. Below is a general equation to calculate the atomic mass of an element based on percent natural abundance and isotopic masses:

$$mass\ atomic = \frac{\sum_n^1 Xi\ Mi}{100}$$

Xi = % of isotope

Mi = masse of isotope

$\sum = 100$

Exampels :

- Bromine has two naturally occurring isotopes: bromine-79 has a mass of 78.9183 u and an abundance of 50.69%, and bromine-81 has a mass of 80.92 u and an abundance of 49.31%. The equation above can be used to solve for the relative atomic mass of bromine:

atomic mass of Br = (50.69 x 78.9183 u+ 49.31 x 80.92 u)/ 100 = 79.91 u

This is the relative atomic number of bromine that is listed on the periodic table. Comparing their isotopic masses of any given element to the relative atomic mass of the element reveals that the A_r is very close to the isotope that occurs most frequently. Thus, the isotope whose isotopic mass is closest to the atomic mass of the element is the isotope that occurs in the greatest abundance.

- Chlorine exists in two isotopic forms, ^{35}Cl and ^{37}Cl. The mass of the ^{35}Cl isotope is 34.97 amu and that of ^{37}Cl is 36.97 amu. The abundances are 75.77% and 24.23%, repectively. Therefore in this case, the weighted average becomes.

Mass atomic of Cl = (75.77 x 34.97u + (24.23 x 36.97u) / 100 = 35.45 amu

The result of this calculation is the atomic mass of chlorine that appears in the periodic table.

Note:

Isobars : are atoms of different chemical elements with equal atomic mass values. The series of elements with 40 Mass numbers serve as a good example; $^{40}_{16}S$, $^{40}_{17}Cl$, $^{40}_{18}Ar$, $^{40}_{19}K$, and $^{40}_{20}Ca$. The nucleus of all the above-mentioned elements contain the same number of particles in the nucleus but contain varying numbers of protons and neutrons.

Isotones : are atoms of different chemical elements with an equal number of neutrons in the atomic nucleus.

For example : boron-12 and carbon-13 nuclei both contain 7 neutrons, and so are isotones . Similarly, $^{36}_{16}S$, $^{37}_{17}Cl$, $^{38}_{18}Ar$, $^{39}_{19}K$, and $^{40}_{20}Ca$ are all isotones of 20 since they all contain 20 neutrons.

IV.3. Uses of isotopes

There are a wide variety of applications of isotopes in nuclear chemistry, medicine, biochemistry, anthropology, paleontology, and geology. Many such uses are based on the phenomenon of radioactivity, shown by some of the isotopes of many of the elements. Such radioactive isotopes are unstable, undergoing spontaneous nuclear decay processes at a rate determined by the **half-life** of the isotope. One example is the use of ^{14}C - the isotope of carbon with six protons and eight neutrons, which has a half-life of 5730 years - as a basis for dating of materials derived from living organisms that are many thousands of years old. This technique, called **radiocarbon dating**, is used widely in geosciences and anthropology.

V. Binding Energy

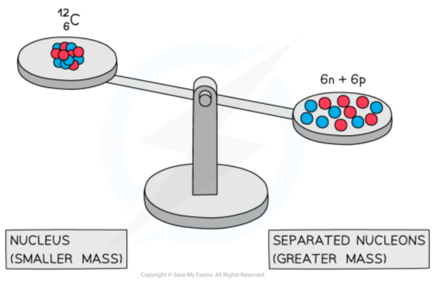
Binding energy can be defined as the work needed to break the nucleus into its components (protons and neutrons) or it is the energy released when the components of the nucleus are combined together.

V.1. Mass Defect

- Experiments into nuclear structure have found that the total mass of a nucleus is **less** than the sum of the masses of its constituent nucleons
- This difference in mass is known as the **mass defect**
- Mass defect is defined as: ***The difference between the mass of a nucleus and the sum of the individual masses of its protons and neutrons***
- The mass defect Δm of a nucleus can be calculated using:

$$\Delta m = Zm_p + (A - Z)m_n - m_{total}$$

- Where:
  - Z = proton number
  - A = nucleon number
  - m_p = mass of a proton (kg)
  - m_n = mass of a neutron (kg)
  - m_{total} = measured mass of the nucleus (kg)



V.2. Binding enrgy

- Binding energy is defined as: ***The energy required to break a nucleus into its constituent protons and neutrons***
- Energy and mass are proportional, so, the total energy of a nucleus is less than the sum of the energies of its constituent nucleons
- The formation of a nucleus from a system of isolated protons and neutrons is therefore an exothermic reaction - meaning that it releases energy
- This can be calculated using the equation:

$$E_b = \Delta mc^2 \text{ ; Where : } \Delta m \text{ in Kg ; c in m/s ; } E_b \text{ in joul}$$

Or

$$E_b = \Delta m\ 931 \text{ ; Where : } \Delta m \text{ in mau; } E_b \text{ in Mev}$$

Since the total binding energy of the nucleus depends on the number of nucleons, a more useful measure of the cohesiveness is the average binding energy Bave.

In nuclear physics, one of the most important experimental quantities is the

- **binding energy per nucleon (E_{bn}) :** which is defined by

$$E_{bn} = E_b / A$$

This quantity is the average energy required to remove an individual nucleon from a nucleus—analogous to the ionization energy of an electron in an atom. If the BEN is relatively large, the nucleus is relatively stable. BEN values are estimated from nuclear scattering experiments.

Calculate the binding energy per nucleon of an $^4He(\alpha particle)$

Determine the total binding energy (E_b)

using the equation : $E_b=(\Delta m)c^2$

where Δm is the mass defect

. The binding energy per nucleon (BEN) is BE divided by A.

Solution

For ^4He

, we have $Z=N=2$. The total binding energy

$E_b=[2mp+2mn]-m(4He)c^2$.

These masses are $m(^4He)$=4.002602 mau, mp=1.007825mau, and mn=1.008665mau

. Thus we have

$E_b=(0.030378u)c^2$.

Noting that $1u=931.5MeV/c2$

, we find

E_b =(0.030378)(931.5$MeV/c2$)$c2$=28.3MeV.

Since $A=4$

, the total binding energy per nucleon is $BEN=7.07 MeV/nucleon$.

Expressing Nuclear Binding Energy as Energy per Mole of Atoms, or as Energy per Nucleon

The energy calculated in the previous example is the nuclear binding energy. However, nuclear binding energy is often expressed as kJ/mol of nuclei or as MeV/nucleon.

- **To convert the energy to kJ/mol of nuclei** we will simply employ the conversion factors for converting joules into kilojoules (1 kJ = 1000 J) and for converting individual particles into moles of particles (Avogadro's Number).
  (8.8387 x 10^{-11} J/nucleus)(1 kJ/1000 J)(6.022 x 10^{23} nuclei/mol) = 5.3227 x 10^{10} kJ/mol of nuclei
- **To convert the binding energy to MeV (megaelectron volts) per nucleon** we will employ the conversion factor for converting joules into MeV (1 MeV = 1.602 x 10^{-13} J) and the number of nucleons (protons and neutrons) which make up the nucleus.

(8.8387 x 10^{-11} J/nucleus)[1 MeV/(1.602 x 10^{-13} J)](1 nucleus/63 nucleons) = 8.758 MeV/nucleon

V.3. .Factors of the syability of he nucleus :

The factors that have an important role in knowing the stability of the nucleus are:

First: The ratio (N/Z) means the ratio of neutrons to protons, Whene :

- N = Z : this ratio will be valued (1) ; the nuclei will be stable
- N > Z : the nuclei will be unstable in a direction Negative beta decay (

B$^-$ electrons).

N < Z : the nuclei will be unstable towards the decay of positive (β^+ positrons).

Second : The odd and even formula for the numbers of protons and neutrons have an important role in the stability of the nucleus ; where if:

- N and Z are a odd number, the nuclei are less stable
- N and Z are a even numbers, the nuclei are more stable
- And while nuclei with an odd number of protons Protons and an even number of neutrons are less stable than nuclei that contain an even number of protons and an odd number of neutrons.

Third: Binding energy per nucleon: The value of the binding energy per nucleon is considered a measure of the cohesion and stability of the nucleus. The higher this value, the coherent and stable the nucleus, and the lower this value, this means that the nucleus is more disintegrating and unstable.