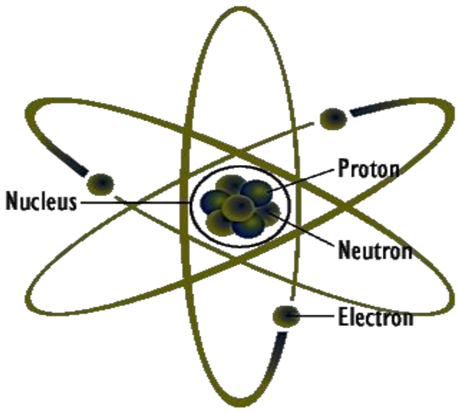
Chapter One: Atomic Structure

❖ Fundamental particles of an atom

*An ***atom*** is the smallest unit quantity of an element that is capable of existence, either alone or in chemical combination with other atoms of the same or another element.

*The fundamental particles (or ***subatomic*** or ***elementary*** particles) of which atoms are composed are the ***Proton***, ***Electron*** and ***Neutron***.



	Proton	Electron	Neutron
Symbol	P^+	e^-	N^0
Charge	+1	-1	0
Rest Mass (*Kg*)	1.673×10^{-27}	9.109×10^{-31}	1.675×10^{-27}
Mass (*amu*)	1	0.0005	1
Location	Inside nucleus	Outside nucleus	Inside nucleus
Discovery	1919, Rutherford	1897, Thomson	1932, Chadwick

* By international agreement, a carbon atom that contains six protons and six neutrons has an atomic weight of exactly 12 ***amu***, so we can define **Atomic Mass Unit** (***amu***) as ***1/12th*** of the mass of a $^{12}_{6}C$ atom so that it has the value 1.660 * 10^{-27} kg.

***** A neutron and a proton have approximately the same mass and, relative to these, an electron has ***negligible*** mass.

Notice**: The mass of an electron is ***much smaller than that of either a ***proton*** or ***neutron***. It takes almost ***2,000 electrons*** to equal the mass of a ***single proton***.

* The charge on a ***proton*** is ***positive***, with (opposite sign), to that on a ***negatively*** charged ***electron***; a ***neutron*** has **no** charge.

*In an atom of any element, there are ***equal*** numbers of ***protons*** and ***electrons*** and so an atom is neutral.

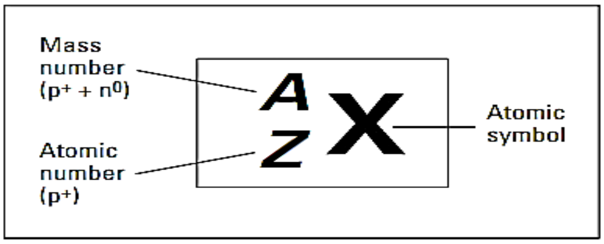
***** The nucleus of an atom consists of ***protons*** and ***neutrons*** (with the exception of ***protium***, the **nucleus** of (**H**) cation has only ***single proton*** without ***neutron***), and is positively charged.

***** The electrons occupy a region of space around the nucleus.

* Nearly all the mass of an atom is concentrated in the nucleus, but the volume of the nucleus is only a tiny fraction of that of the atom.

* The radius of the nucleus is about **10^{-15} m** (1 femtometer = 10^{-15} m) while the atom itself is about **10^{5} times** larger than this.

❖ Atomic number, mass number and isotopes



* ***Atomic number***, ***Z***, which is equal to the number of ***protons*** (***p^+***).

* In electrically neutral atom, ***Z*** also equals the number of ***electrons*** (**e**$^-$).

* ***<u>Mass number</u>*** (or ***Atomic Weight***), ***A***, is the number of ***protons*** (***p***$^+$) and ***neutrons*** (**n**^{0}) in the ***nucleus***.

* So number of ***neutron***

$$n^0 = \textit{Atomic Weight (A) - Atomic number (Z)}$$

* **<u>Ex</u>:- Fe** nucleus which has **(*26 p*$^+$)** and **(*30 n*^{0})**, is denoted as $^{56}_{26}$**Fe**.

****<u>Isotopes</u>:-*** Are atoms of the same element that differ only in the number of neutrons (mass number or atomic weight). Some isotopes occur naturally while others may be produced artificially.

*** <u>Isotopes</u>: -** (Atoms of the same element that have an ***identical*** number of ***protons*** (***same atomic number***) but ***different*** numbers of ***neutrons*** (***different mass numbers***). The figure below shows the three hydrogen isotopes, *Hydrogen* or (**Protium**) which has (1p$^+$, 0 n^0), ***Deuterium*** which has (1p$^+$, 1n^0) and ***Tritium*** which has (1p$^+$, 2n^0).

$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
(a) Hydrogen	(b) Deuterium	(c) Tritium
H – 1	H – 2	H – 3

Hydrogen Isotopes

* **<u>Isotopes</u>** have the same position in the periodic table, the same chemical properties, and the same atomic charge.

* **<u>Isobars</u>**:- Are atoms of different elements that have the same mas number, but they have different atomic numbers. For instance Cl-37 and Ar-37 have the same mass number.

* Table below shows the average isotopic abundances of some elements.

Hydrogen (H)	Carbon (C)	Nitrogen (N)	Oxygen (O)
^{1}H – 99.984%	**^{12}C – 98.89%**	**^{14}N – 99.64%**	**^{16}O – 99.763%**
^{2}D – 0.0156%	**^{13}C – 1.11%**	**^{15}N – 0.36%**	**^{17}O – 0.0375%**
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*__Note__:- Since the atomic number is constant for a given element, isotopes are often distinguished only by stating the atomic masses, e.g. ^{12}C and ^{13}C.

❖ ***Relative atomic mass (RAM):-*** The relative atomic mass of an atom is the weighted mean of the mass numbers of its isotopes.

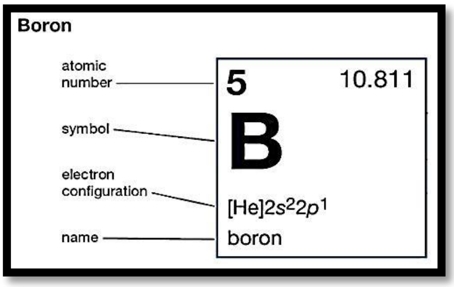
$$RAM = \frac{[(\text{abundance of 1st isotope\%}) * (\text{mass of isotope})] + [(\text{abundance of 2nd isotope\%}) * (\text{mass of isotope})] + (\text{etc.})}{100}$$

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

$$\textbf{Relative atomic mass (RAM)} = \frac{[(19.9\%)\ (10.013)] + [(80.1\%)\ (11.009)]}{100}$$

$$= 10.811$$

(Note: - that this is the value of atomic mass given on the periodic table).



H.W. Calculate RAM for naturally occurring Mg if the isotope distribution is 78.99% ^{24}Mg, 10.00% ^{25}Mg and 11.01% ^{26}Mg; accurate masses are 23.99, 24.99 and 25.98. **[Ans. 24.31]**

H.W. Calculate the value of RAM for naturally occurring chlorine if the distribution of isotopes is 75.77% $^{35}_{17}Cl$ and 24.23% $^{37}_{17}Cl$. Accurate masses for ^{35}Cl and ^{37}Cl are 34.97 and 36.97. **[Ans. 35:45]**