The elements, compounds, and atoms

- All substances are made up of atoms, which are the smallest possible unit of an element.
- Chemical symbols, on the other hand, stand in for individual atoms of an element.
- Chemical symbols represent an atom of an element e.g. Na represents an atom of sodium
- Chemical reactions invariably result in the creation of one or more new substances, and they frequently entail an energy shift.
- Compounds are composed of two or more elements combined chemically in predetermined amounts; they can be represented by formulae using the atom symbols from which they were formed, for example.
- One atom of hydrogen and one atom of chlorine make up the compound HCl. Chemical reactions are the only way to split compounds into their constituent parts.

Mixture

- A mixture is made up of two or more substances that aren't joined chemically. Every material in the mixture has the same chemical characteristics. (This is not the same as a compound.)
- Can be separated *via* chromatography, simple distillation, fractional distillation, filtration, and crystallisation. Since these are physical processes, no new compounds are created or chemical reactions take place.

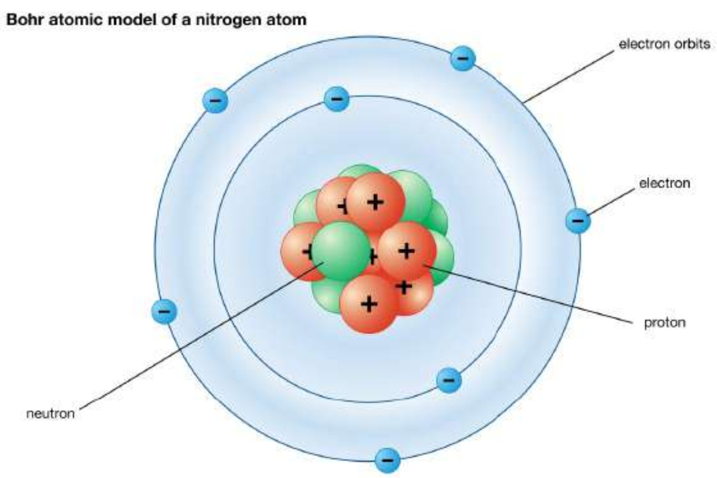
The evolution of the model of the atom

- Initially, it was believed that atoms were tiny spheres that could not be divided. Later, the discovery of the electron led to the development of the plum pudding model, which states that an atom is a ball of positive charge with embedded negative electrons.
- Afterward, an alpha particle scattering experiment revealed that an atom's mass was concentrated at its nucleus and that the nucleus was charged.

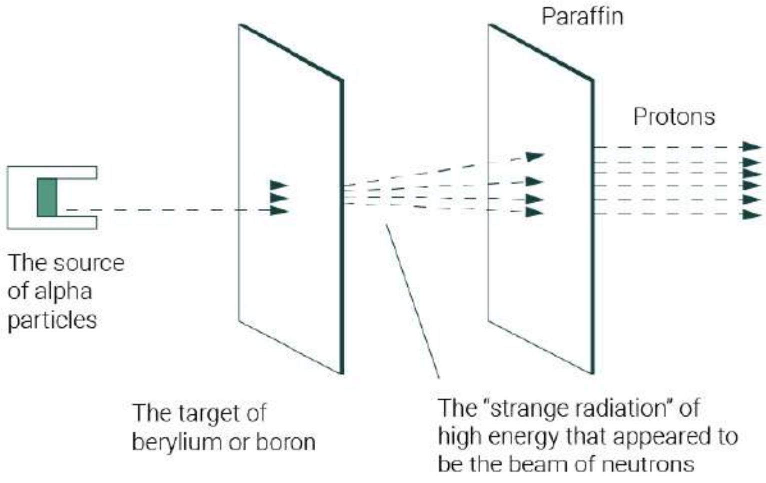
Experiment with scattering:

- A beam of alpha particles was aimed at very thin gold foil and their passage through was detected Some of the alpha particles emerged from the foil at different angles, and some even came straight back the positively charged alpha particles were being repelled and deflected by a small concentration of positive charge in the atom (nucleus).

- Neil Bohr: suggested electrons orbit the nucleus at specific distances (supported by experimental data) Later experiments: positive charge of any nucleus could be subdivided into a whole number of smaller particles, each particle having the same amount of positive charge (protons).



- James Chadwick's work: provided the evidence to show the existence of neutrons within the nucleus (had been an accepted scientific idea for about 20 years already)



The proportional electrical charges of subatomic particles

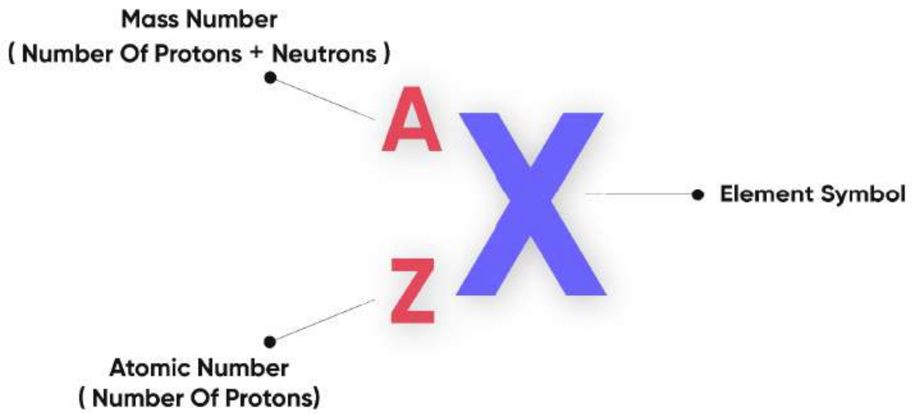
- An element's atomic number is its total number of protons; all atoms of a given element have the same number of protons.
- The quantity of protons varies amongst elements of atoms.
- The relative charge of a particle is +1 for a proton, 0 for a neutron, and -1 for an electron.
- Since the total charge of an atom is zero, the number of protons equals the number of electrons.

Relative electrical charges of subatomic particles

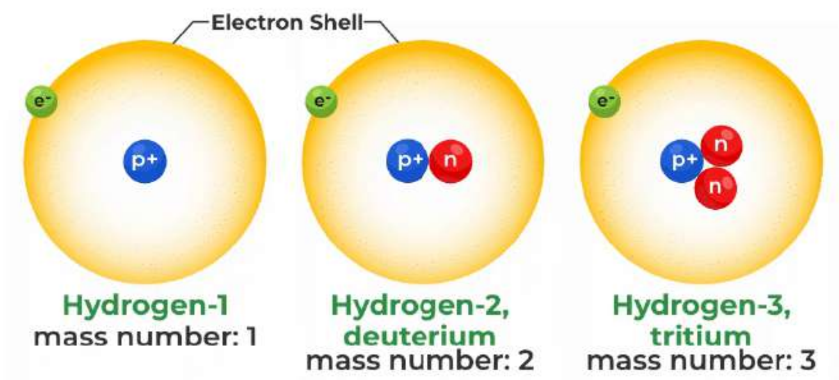
- Atomic number: the number of protons in an atom of an element.
- All atoms of a particular element have the same number of protons.
- Atoms of different elements have different numbers of protons.
- The proton has relative charge as +1, neutron has relative charge as 0, and electron has -1
- An atom has an overall charge of 0, so number of protons = number of electrons

Size and mass of atoms

- Atoms are very small (radius of about 0.1 nm) and the radius of a nucleus is less than 1/10,000 of that of the atom, though it holds almost all of the mass particle.
- The Relative mass of proton is 1, for neutron it is 1 and for electron it is very small
- Mass number: the sum of the protons and neutrons in an atom.



- Isotopes: atoms of the same element with different numbers of neutrons.



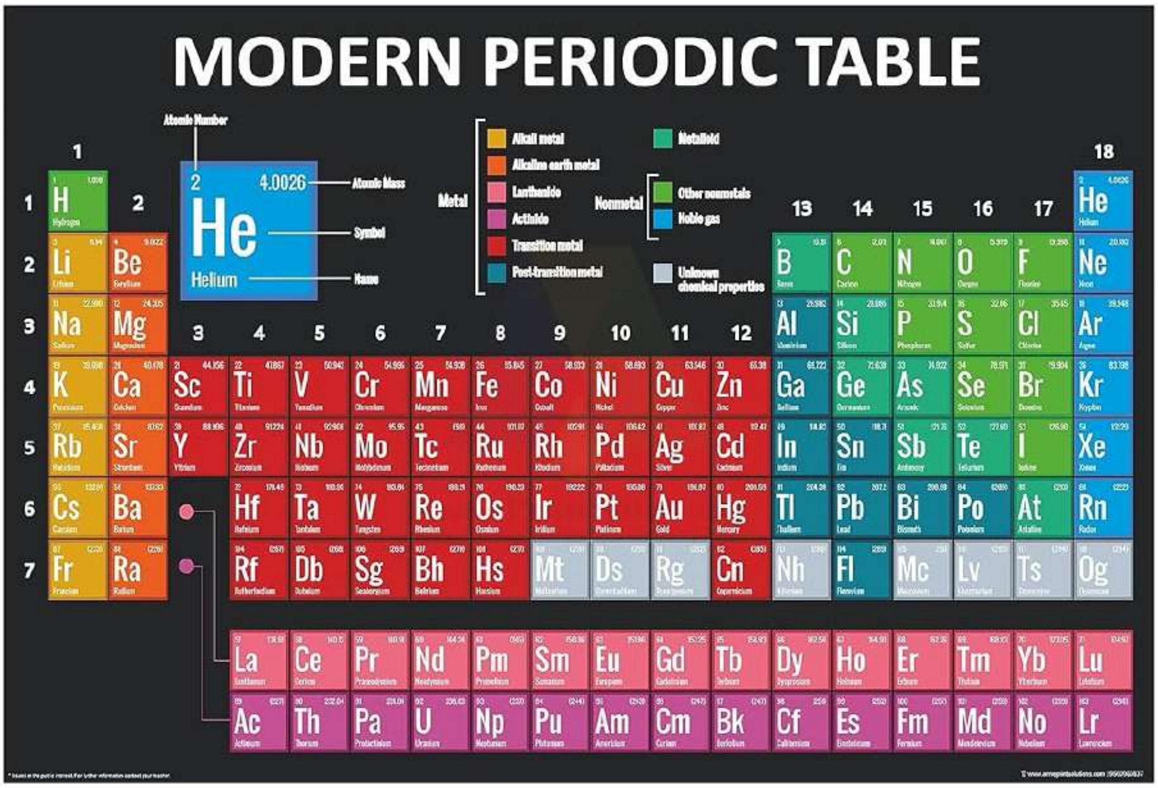
Relative atomic mass

- Relative atomic mass: an average value that takes account of the abundance of the isotopes of the element.
- Example question: carbon has 2 isotopes: carbon-14 with abundance 20% and carbon-12 with abundance 80%.

- Calculate the relative atomic mass of carbon. To calculate it: ((isotope 1 mass x abundance) + (isotope 2 mass x abundance)) ÷ 100 For this question: ((14 x 20) + (12 x 80)) ÷ 100 = 1240 ÷ 100 = 12.4

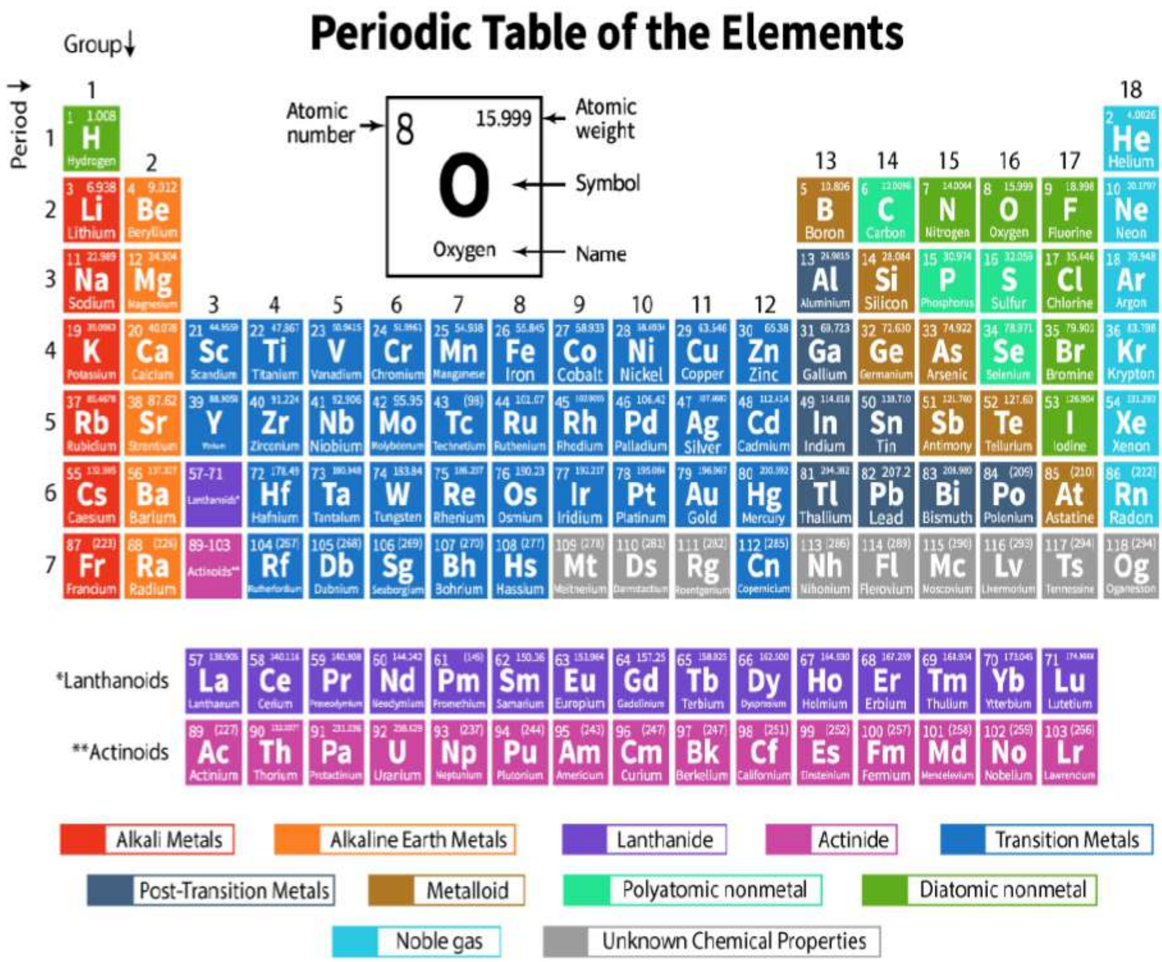
- The electrons in the electronic structure inhabit the lowest energy levels, which are the shells nearest to the central nucleus.
- An atom's electronic structure indicates the number of electrons in each shell, such as in the case of sodium: two electrons within the shell 1 in shell 2 (nearest to the nucleus), 8 in shell 3 Electronic structure= 2,8,1 or
- The periodic table is based on atomic masses and related features. The atomic masses in each row go up and to the right. Each column contains a group of elements with similar chemical behaviour.

In the modern periodic table, each box contains four data, as shown in Figure.



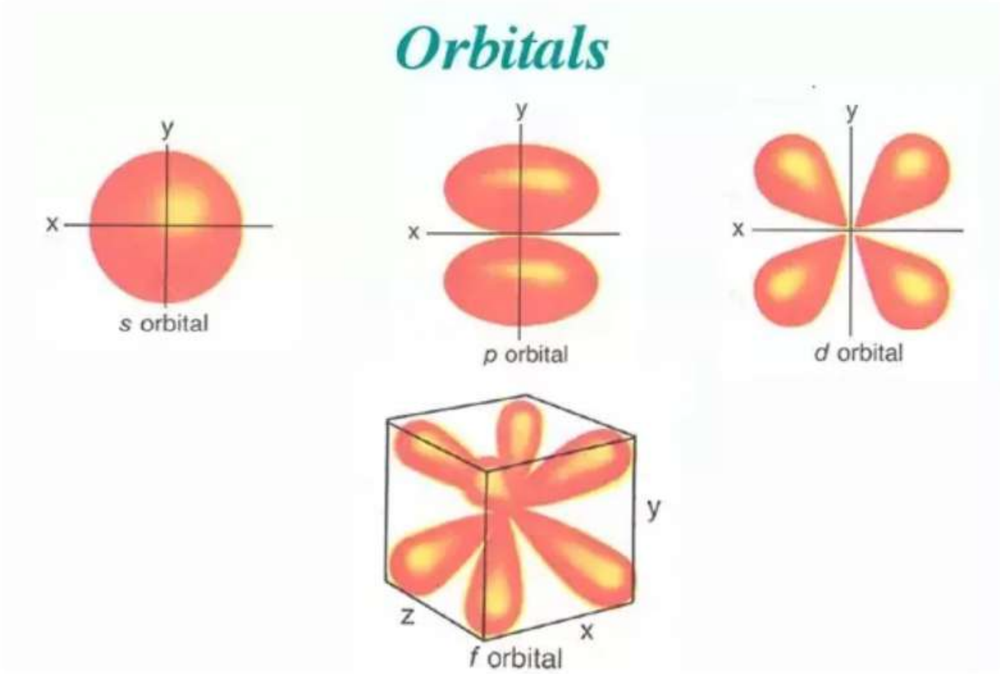
Besides the element name and symbol, the atomic mass is at the bottom, and the atomic number is at the top. The elements are arranged in order of increasing atomic number in horizontal rows called periods. The atomic number, which appears above each element symbol, represents the meaningful order in the periodic table. When an element is referred to by an integer, this number means the atomic number, not the atomic mass. Thus, element 27 is cobalt (whose atomic number is 27), not aluminium (whose atomic mass is 27).

The periodic table displays the pattern of properties of the elements. The lightest are at the top of the chart; the atomic masses increase toward the bottom of the chart. The elements to the upper right, above a diagonal line from aluminium (13) to polonium (84), are nonmetals, about half of which exist as gases under normal laboratory conditions. All the elements in the middle and left of the table are metals, except gaseous hydrogen. Most of the metals are shiny, deformable solids, but mercury has such a low melting point that it is a liquid at room temperature. All the metals have high conductivities for heat and electricity. Many simple chemical compounds are formed from a metal reacting with a nonmetal.



- Modern periodic law is used to create the long version of the periodic table, commonly known as the Modem Periodic Table.
- The elements have been organised in this table in order of increasing atomic numbers.

ELECTRONIC ARRANGEMENTS OF THE ELEMENTS IN THE s, p, d, and f blocks



S-block

- The final electron in these elements reaches the s-orbitals.
- S-block components typically have an outer shell electronic configuration of ns1-2, where n = 2–7.
- These metals have lower melting and boiling temperatures and are softer.
- They are electropositive (energies) and have low ionisation enthalpies.
- They quickly lose their outermost (valence) electrons, becoming +2 ions (for alkaline earth metals) and +1 ions (for alkali metals).
- metals that are extremely reactive. As we move through the group, the metallic quality and responsiveness become increasingly apparent. Pure ones are rare in nature because of their intense responsiveness.
- Except for beryllium, all compounds containing s-block elements are ionic.

p-block

- A general electrical setup is ns2np1-6.
- The d-block element compounds are primarily covalent in nature.
- They exist in different levels of oxidation.
- From left to right, the components get more non-metallic as the era goes on.
- Component responsiveness tends to decline within a group.
- A noble gas element with a closed valence shell (ns2np6) configuration is present at the end of each cycle.
- The metallic aspect becomes more and more apparent as we move through the group.

d-block

- A common electrical layout is (n-1)d1-10ns1-2.
- All of them are metals with high boiling and melting temperatures.
- D-block element compounds are often paramagnetic in nature.
- The majority of the time, they produce coloured ions with a variety of valences (oxidation states).
- Because the d-block elements have partially filled d-orbitals in either their ground state or any of their oxidation states, they are referred to as transition elements.
- Compounds containing the d-block elements typically make up catalysts.

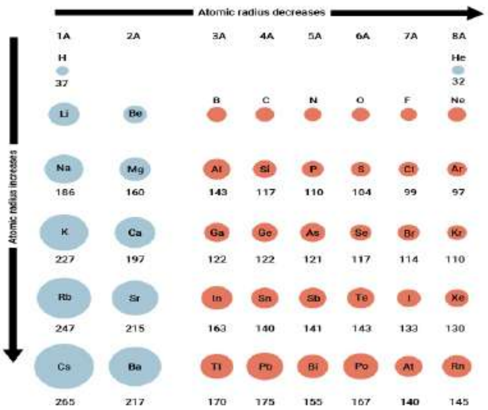
f-block

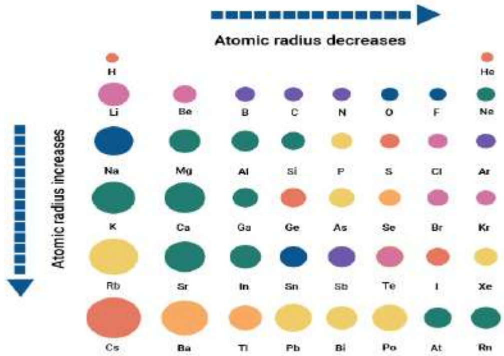
- The overall electrical configuration is n-2f1-14((n2)f(014)(n1)d(01)ns2.
- The f-block inner transition elements are called inner transition elements because their electrons are filled in (n – 2) f subshells, which is one inner subshell, as opposed to the (n – 1) d subshells found in d-block transition elements.
- The Lanthanoids Ce (Z = 58) – Lu (Z = 71) and the actinoids Th (Z = 90) – Lr (Z = 103) comprise the two rows of elements at the bottom of the Periodic Table.
- They are made entirely of metal. Many of the components in each series are similar to one another.
- The actinoid series contains a significant amount of radioactive materials.

Periodic Trends in Properties of Elements

Changes in Physical Properties

Atomic Radii: The unit of measurement for the separation between the core of the nucleus and the outermost electron-containing shell is the atomic radius. Different atomic radii are used for different elements based on whether they are metals or non-metals. They are as follows





(a) Covalent Radius: The half-distance between two atoms' nuclei that are joined by a single, exclusively covalent connection.

(b) Ionic Radius: The effective distance that an ion has an impact on an ionic bond from its nucleus.

(c) Van der Waals Radius: Noble gas atoms are held together by the weak van der Waals forces of attraction. Half of the distance between the nuclei of noble gas atoms is the van der Waals radius.

(d) Metallic Radius: This term refers to the half of the internuclear distance between two adjacent metal ions in the metallic lattice.

Radius of Ions

(a) Ionic radii in ionic crystals can be calculated from the distances between cations and anions.
(b) Ionic radii of elements generally exhibit the same pattern as atomic radii.
(c) When an electron is taken out of an atom, a cation is produced. The radius of the cation is always less than that of the atom.
(d) Gaining an electron results in the formation of an anion. The radius of the anion is always larger than the radius of the atom.
(e) Isoelectronic Species: Atoms and ions with the same electron count are referred to as isoelectronic species. For instance, the number of electrons in O2-, F–, Na+, and Mg2+ is the same. Their radii would differ due to their different nuclear charges.

The quality of electrons

(a) An atom's ability to draw other electrons from a chemical combination to itself is measured qualitatively by its electronegativity.
(b) Unlike ionisation and electron gain enthalpy, it cannot be measured.
(c) Nonetheless, a number of numerical scales, such as the Pauling, Milliken-Jaffe, and Allred Kochow scales, have been developed for the electronegativity of various elements.
(d) The electronegativity of each given element varies according to the element it is associated with.

Periodic Patterns in Chemical Characteristics across Time

(i) Metal properties: A decrease over a period of time with a maximum on the far left (alkali metals).
(ii) Non-metallic nature: Develops over time to become increasingly non-metallic. (Going from left to right)
(iii) Oxides' fundamental nature: They diminish with time from left to right.
(iv) Oxides' acidic nature: Rises over time from left to right.

Distinctions When Declining a Group from the Top to the Bottom

(i) It exudes a metallic vibe. When atomic size increases, the ionisation energy of the elements in a group falls as well, causing a decline in ionisation energy from top to bottom.

(ii) Non-metallic characteristics. decreases a group's overall size. Elements in a group lose electronegativity from top to bottom.

The essential nature of oxides

(iii). Since the metallicity or electropositivity of an element grows from top to bottom in a group, the essential nature of oxidises naturally.

(iv) Oxides' acidic nature. In general, the non-metallic property of an element reduces as it moves down the group from top to bottom.

(v) Reactivity of metals. A group's size typically increases as its membership expands. due to the growing possibility of electron loss.

(vi) Reactivity without metals. Non-metals are more reactive the higher their electro-negativity. As the electronegativity of non-metals in a group decreases from top to bottom, so does their reactivity.