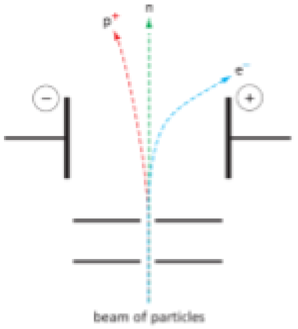
# UNIT 1 - ATOMIC STRUCTURE

Atom - the smallest part of an element that can take part in a chemical change.

Atoms are made up of subatomic particles. These include protons, neutrons, and electrons. Each subatomic particle has a fixed mass and charge.

| | Relative mass | Relative charge | |
|---|---|---|---|
| **Proton** | 1 | +1 | $1.672 \times 10^{-24}$ g |
| **Neutron** | 1 | nil | $1.675 \times 10^{-24}$ g |
| **electron** | $1/1836$ | −1 | $9.109 \times 10^{-28}$ g |

The nucleus is a tiny area in the middle of every atom and is where almost all of the mass is concentrated. The nucleus is made up of particles called nucleons, which include protons and neutrons. Different elements' atoms have varying quantities of protons. Electrons orbit the nucleus in orbitals. To determine the weight and charge of subatomic particles, you can pass them through electrically charged plates and measure the angle of deflection.



The heated wire emits electrons, which are drawn to two positively charged metal plates by their positive charges. The electrons produce a beam as they move through the metal plates. A pinpoint of light is formed on the screen when the electron beam strikes it. The electrons are deflected when an electric field is introduced across this beam. It is clear from the ease with which the electrons are drawn to the positively charged anode and the ease with which they are redirected by an electric field that: electrons have a negative charge; electrons have a very low mass.

Angle of Deflection $\propto \frac{q}{m}$

q → charge

m → mass.

$\measuredangle = k\,q/m$

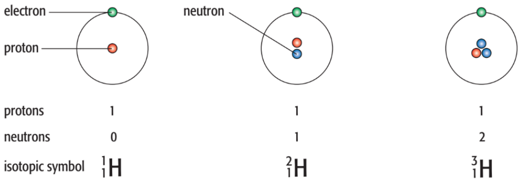
## Isotopes

Isotopes are atoms with:

- The same number of protons and electrons but a different number of neutrons;
- The same atomic number but a different mass number.

The chemical properties of isotopes are identical, whereas, the physical properties (such as density and mass) can differ. When writing generally about isotopes, chemists also name them by leaving out the proton number and placing the mass number after the name. For example, the isotopes of hydrogen can be called hydrogen-1, hydrogen-2, and hydrogen-3.

An example is the following isotopes of hydrogen:

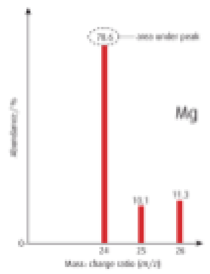


## Relative Masses

Relative Atomic Mass (Ar) - the weighted average mass of the naturally occurring isotopes of the element relative to one-twelfth the mass of a Carbon-12 isotope.

Relative Isotopic Mass - the mass of an atom of the isotope of the element relative to one-twelfth the mass of an atom of the isotope Carbon-12.

## Mass Spectrometry



The mass spectrum is a plot of relative abundance (of each isotope) versus m/e or the mass number.

The height of each peak indicates the relative abundance of the respective isotope.

A mass spectrometer is an analytical machine used to calculate the relative atomic mass of an atom by comparing it with the mass of a 12C atom.

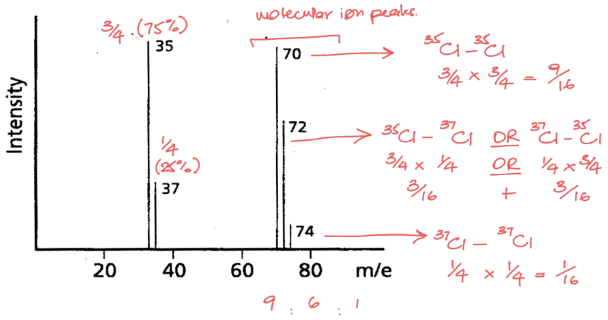
- It is also used to calculate the relative abundance of different isotopes of any element.

To calculate relative isotopic mass from a mass spectrometer graph like this, take the percentage abundance and multiply it by the corresponding m/e, and divide it by 100:

$$(78.6 \times 24) + (10.1 \times 25) + (11.3 \times 26) = 2432.7$$
$$2432.7 / 100 = 24.33$$

Molecular Ions

Molecules also give peaks in their mass spectrum. For example, the mass spectrum of chlorine gaseous molecule includes both peaks of its isotopes and peaks of its molecular form (molecular ions).



Electrons

Electrons are arranged into different levels of energy. They are arranged outside the nucleus in energy levels or quantum shells. These energy levels or principal quantum shells (n) are numbered according to how far they are from the nucleus. The electrons further away from the nucleus have more energy and are less tightly held.

Ionisation Energy

By firing high-speed electrons at atoms, scientists can work out how much energy has to be supplied to form an ion by knocking out one electron.

1st ionization energy - the energy needed to remove 1 mole of the electron from 1 mole of its gaseous atom, under standard conditions.

The general symbol for Ionisation Energy is - $\Delta H_{I.E}$ or $1^{st}$I.E and its units are kJ $mol^{-1}$.
If a second electron is removed from each ion in a mole of gaseous 1+ ions, we call it the second ionization energy, $IE_2$. If a third electron is removed from each ion in a mole of gaseous 2+ ions, we call it the third ionization energy, $IE_3$, and so on.

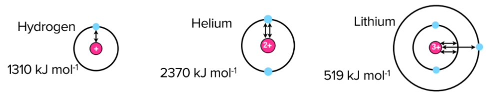
If we continue to remove electrons from an atom until only the nucleus is left, the sequence of ionization energies is called successive ionization energies.

- The successive ionization energies rise for each element. This is due to the fact that as each electron is taken out, the ion's net positive charge increases. The attractive force between the positively charged protons in the nucleus and the remaining negatively charged electrons increases as each electron is eliminated. Therefore, more energy is needed to overcome these attractive forces.

- There is a big difference between some successive ionization energies. For nitrogen, this occurs between the 5th and 6th ionization energies. For sodium, the first big difference occurs between the 1st and 2nd ionization energies. These large changes indicate that for the second of these two ionization energies, the electron is being removed from a principal quantum shell closer to the nucleus.

*Factors affecting first ionization energy:*

The value of the first Ionisation Energy depends on the electronic structure.



The further the valence electron is from the nucleus, the easier it is to remove, and so the smaller the ionization energy.

Other factors include:

**Nuclear Charge**
As the nuclear charge increases the attractive force would increase and hence more energy would be needed to overcome the increasing nuclear attraction.

**Atomic Radius**
As the atomic radius increases, the outermost electron would be further away from the nucleus experiencing a weaker attractive force. Hence the value decreases with increasing size. Attraction decreases very rapidly with distance. An electron close to the nucleus will be much more strongly attracted than one further away.

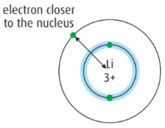
**Shielding effect/Screening**

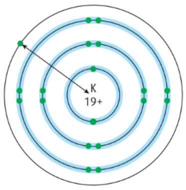
Electrons in full inner shells repel electrons in outer shells. Full inner shells of electrons prevent the full nuclear charge from being felt by the outer electrons. The inner electron orbitals effectively screen or shield the outermost electrons from the nucleus, due to which the ionization energy decreases. The lessening of the pull of the nucleus by the inner electrons is known as shielding. It is due to the shielding effect that the electrons in the outer shell are attracted by the effective nuclear charge which is less than the full charge on the nucleus. ONLY S ORBITALS PROVIDE SHIELDING.

**The nature of the sub-level**

Electrons in the s orbital is closer to the nucleus as compared to the electrons in the p orbital of the same shell. Hence the s electrons are the more firmly held and require a greater amount of energy to be removed as compared to the p electrons which are slightly further. Also, the p orbital of the same energy level will experience the shielding effect of the s electrons from the same shell in addition to previous s orbitals.

*Ionization Energy down a group*





Ionization energy decreases down a group. Despite an increased nuclear charge the outer electron is easier to remove. This is due to greater distance from the nucleus and increased shielding. Down the group, as the number of shells increase, the outer electrons are farther away from the nucleus.

Also, down the group each successive element contains more electrons in the shells between the nucleus and the outermost electrons which increases the shielding effect. This increased shielding causes the outermost electrons to be held less tightly to the nucleus. The increased shielding effect and increased atomic radius/size outweigh the increase in nuclear charge. Therefore ionization energy decreases.

*Ionization energy across a period*

**General Trend: Ionisation energy increases across a Period**

Across a period, the number of protons and number of electrons increase by one each. The additional proton increases the nuclear charge. The additional electron is added to the same outer shell in each of the elements.

- A higher nuclear charge more strongly attracts the outer electrons in the same shell, but the electron-shielding effect from inner-level electrons remains the same. Thus, more energy is required to remove an electron because the attractive force on them is higher.

<u>Shells and sub-levels</u>

Each energy level (shell) consists of a number of sub-levels (sub-shells) labeled s, p, d, or f.

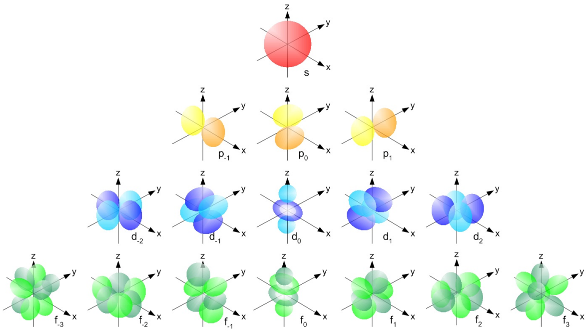
Each sub-shell can hold a certain maximum number of electrons.

s - 2 electrons
p - 6 electrons
d - 10 electrons
f - 14 electrons

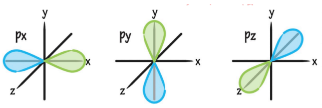
| Energy Level | Number of sub-levels | Name of sub-levels |
|---|---|---|
| 1 | 1 | s |
| 2 | 2 | s, p |
| 3 | 3 | s, p, d |
| 4 | 4 | s, p, d, f |

An atomic orbital is a region of space around the nucleus of an atom that can be occupied by one or two electrons only. Each sub-shell contains a fixed number of orbitals that contain electrons.

s - 1 orbitals
p - 3 orbitals
d - 5 orbitals
f - 7 orbitals



The s orbital is at the top and is a sphere shape.



The p orbitals are the yellow ones on the second level of the pyramid. They are two dumbbell shapes and can be plotted on the x-axis, y-axis, or z-axis. To draw these in an exam, you can draw them as sideways 8 shapes.

The d and f shapes are not required for this syllabus