Subject: Chemistry	Topic: Atomic Structure 3.1.1	Year Group: 12

Subatomic particles

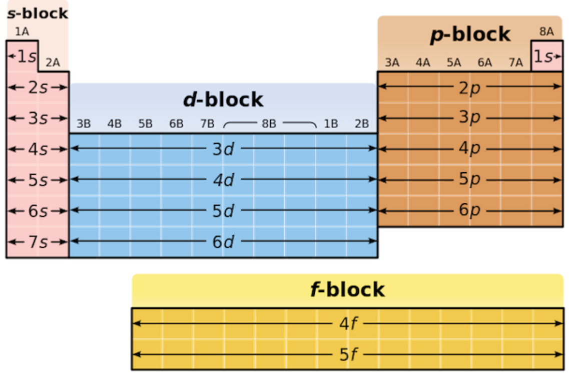
1

Particle	Relative mass	Relative Charge	Location in the atom
Proton	1	-1	Nucleus
Neutron	1	0	Nucleus
Electron	1/1840	+1	Energy levels

2 The number of electrons in an atom is the same as the number of protons (the atomic number), as the negatively charged electrons cancel out the positive charge from the protons.

Blocks

1



2 The periodic table is arranged into the s, p, d and f blocks. All of the elements in the s block have their highest-energy electron in an s orbital, those in the p block have their highest-energy electron in a p orbital, and so on.

Electronic configuration

1 Energy levels (shells): The first energy level (n=1) is the one closest to the nucleus and has the lowest energy, with n=2 being further from the nucleus and having more energy etc. This n is called the principle quantum number. n=1 is made of a single s orbital (called 1s), n=2 is made of an s and a p orbital (2s and 2p), n=3 is made of 3s, 3p and 3d, and n=4 is made of 4s, 4p, 4d and 4f.

2 Sub-shells: These are orbitals or combinations of the same type of orbital. There are 4 types of sub-shell, s, p , d and f, however we only study the first three. Each s sub-shell comprises one s orbital; each p sub-shell comprises 3 p orbitals (px, py and pz); each d sub-shell has 5 d orbitals and each f sub-shell has 7 f orbitals.

3 Orbitals: An orbital is a region of space that can hold up to two electrons (which must be of opposite spin).

4 Electrons fill the lowest energy orbitals first, before filling higher energy orbitals. When filling p and d orbitals, electrons fill singly before pairing up. Within an orbital, if the first electron is spin up, then the second will be spin down.

5 Electrons fill in the following order:
1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p…
Note that the 4s orbital fills before the 3d.

6 The electronic configuration is shown by writing each subshell in order, along with the number of electrons in each (in superscript). For example, Bromine, with 35 electrons is:
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

Key Vocabulary

1	Atomic number (Z)	The number of protons in an atom.
2	Mass number (A)	The total number of protons and neutrons in an atom.
3	Isotopes	Atoms with the same number of protons but different number of neutrons.
4	Subatomic Particles	Particles that make up an atom—i.e. protons, neutrons and electrons
5	Ionisation energy	The energy required to remove an electron from each atom in a mole of gaseous atoms.
6	Second ionisation energy	The energy required to remove an electron from each ion in a mole of gaseous 1+ ions.
7	Isoelectronic	Having the same electronic configuration.

Electronic configuration of ions

1 Ions are formed when atoms lose or gain electrons in order to a more stable electronic configuration i.e. that of a noble gas e.g. [Ne], [Ar] etc. The highest energy electrons are removed first, and these are usually the last electrons that were added.

2 However, when ions are formed from transition elements, the first two electrons are removed from the 4s/5s orbital. This is because when electrons fill the 4s orbital, their energy increases to a value above that of the 3d orbitals, due to electron-electron repulsion.

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Isotopes

1	Isotopes have the same atomic number but different mass number, they have the same number of protons but different number of neutrons.
2	The number of protons defines an element e.g. any atom with 11 protons is a sodium atom. While changing the number of electrons or neutrons does not change the type of atom, changing the number of protons does.

Time-of-flight mass spectroscopy (TOFMS)

1	TOFMS allows us to find the Ar of an element or Mr of a compound. The process involves the following stages: Electrospray ionisation: a high voltage is applied at the tip of a thin capillary containing the substance dissolved in a polar solvent. The solvent evaporates, producing gaseous charged ions. Acceleration: the ions are accelerated by an electric field, with all ions receiving the same kinetic energy. The energy the particles have can be calculated with the equation k.e. = 1/2 mv2 where m is the mass of the atom (calculated by dividing the mass number by Avogadro's constant) and v is the velocity of the particle in m/s. The velocity can be found with the equation v = d/t where d is the distance in metres (the length of the flight tube) and t is the time in seconds. Ion drift: the ions move through a flight tube. Since all particles have the same kinetic energy, heavier particles move slower than lighter ones, therefore taking longer to pass through the tube. Detection: the detector records the different flight times of the ions. When the ion hits the charged detector plate, a current is generated. The greater the current, the greater the number of ions hitting at that point in time, therefore the greater the abundance.

Finding Ar and Mr

1	TOFMS can be used to calculate the relative atomic mass in one of two ways. 1. If the % abundance of each isotope is given, multiply the mass number of each isotope by its % abundance. Then take the sum of these values for all the isotopes, and then divide by 100. 2. If the relative abundance of each isotope is given, multiply the mass number of each isotope by its relative abundance. Then take the sum of these values for all the isotopes, and then divide by the sum of the relative abundances.
2	The mass spectrum of a compound has two peaks of interest: the base peak, which is the tallest peak and represents the most stable ion formed from the fragmentation of the molecule; and the molecular ion peak which is the last major peak found on the spectrum, and whose m/z value is the Mr of the molecule. The species that causes this peak is always shown as an ion e.g. $[C_2H_5OH]^+$.

Trends in atomic radius

1	Atomic radius increases down a group because there are more electrons which occupy higher-energy shells that are further from the nucleus. Atomic radius decreases across a period because additional electrons are placed in the same shell, and nuclear charge increases, which attracts the electrons more strongly, making them closer to the nucleus.

Ionisation energy

1	Ionisation energy is essentially a measure of the energy required to remove an electron from an atom. The higher the energy, the greater the attraction between the negative electron and the positive nucleus, and therefore the more difficult it is to remove.

Trends in ionisation energy

1	First ionisation energy decreases down a group because atomic radius increases and shield by inner electrons increases so less energy is required to remove the electron. First ionisation energy increases across a period because atomic radius decreases, nuclear charge increases and shielding by inner electrons stays the same so more energy is required to remove the electron.
2	Group 3 and 6 elements have lower first ionisation energies than expected because: The electron removed from a Group 3 element is in a p subshell, which is further from the nucleus than an s electron. Also, those inner s electrons provide additional shielding, so the p electron is less attracted to the nucleus and therefore easier to remove.
3	Group 6 atoms have a p4 arrangement. Repulsion of two electrons in the p orbital reduces the amount of shielding, so the p electron is less attracted to the nucleus and so easier to remove.