CIE Chemistry A Level

2 : Atomic Structure

Notes

Particles in the atom

Protons, neutrons and electrons

Particle	Relative mass	Relative charge
Proton	1	1+
Neutron	1	0
Electron	1/1836	1-

When subatomic particles are passed between two oppositely charged plates the protons will be deflected on a curved path towards the negative plate because they are positive. Whereas the electrons will be deflected on a curved path towards the positive plate because they are negative. Neutrons will continue on a straight path because they have no charge.

If the particles have the **same energy** when they are passed between two charged plates, the amount of deflection of the beam of protons and the beam of electrons will be exactly the same in opposite directions. If the particles have the **same speed** when they are passed between two charged plates, the lighter electrons will be deflected more strongly than the protons.[1]

Mass and charge in an atom

The **mass is concentrated at the nucleus** of an atom. The positively charged protons are in the nucleus while the negatively charged electrons orbit the nucleus in shells.

Atomic and mass numbers

The **atomic number** is equal to the **number of protons** in an atom. Atoms of the same element have the same atomic number and number of protons. An atom has **no charge** overall so the **number of electrons must equal to the number of protons** to cancel the charges. An ion is charged so the number of electrons is equal to the atomic number plus or minus the number of electrons gained or lost (a positive charge indicates the number of electrons that have been lost and a negative charge indicates the number of electrons that have been gained).

The **mass number** is equal to the total **number of protons and neutrons** in an atom.

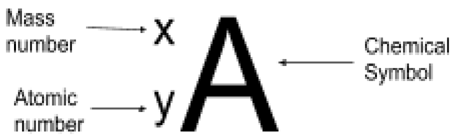
[1] This explanation is based on 'Chemguide' notes by Jim Clark 2010.

The nucleus of the atom

The nucleus contains protons and neutrons (known collectively as nucleons).

Isotopes

Isotopes are atoms of an element with the same number of protons and electrons but a **different number of neutrons**. Isotopes of the same element have **different mass numbers** because they have different numbers of neutrons. Below is how isotopes can be represented:



Electrons

Orbitals

An **orbital** is a region of space where **up to 2 electrons** can be found. A **principal quantum number** (n) represents the shell that the electrons occupy. The larger the principal quantum number, the higher the energy and the further the shell is from the nucleus. The types of orbitals are:

- s: spherical shape, one s orbital in each shell from n = 1 upwards (a total of two s electrons per shell), lowest energy.
- p: dumb-bell shape, three p orbitals in each shell from n = 2 upwards (a total of six p electrons per shell), higher energy than s.
- d: five d orbitals in each shell from n = 3 upwards (a total of 10 d electrons per shell), higher energy than p.

Electrons fill orbitals from lowest energy to highest. 1s is filled first followed by 2s, 2p, 3s, etc. 4s has a lower energy than 3d so **4s is filled before 3d**. Before the electrons start pairing, a subshell must be filled with unpaired electrons. A **subshell** is a specific type of orbitals in a shell (e.g. the p subshell contains 3 p orbitals).

There are two main **exceptions** to electron configuration. A completely full or half full d sublevel is more stable than a partially filled d sublevel, so an electron from the 4s orbital is excited to the 3d orbital.

Chromium: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$

Copper: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

When the ions of copper and chromium are formed, the electrons are removed from the 4s orbital first.

Electron configuration

Electron configurations show the **number of electrons** and **types of orbitals** in each energy level.

Element/ ion	Electron configuration	Comments
B	$1s^22s^22p^1$	2 energy levels, 3 electrons in outer shell, 5 electrons in total
Ne	$1s^22s^22p^6$	2 energy levels, 8 electrons in outer shell, 10 electrons in total
Cl	$1s^22s^22p^63s^23p^5$	3 energy levels, 7 electrons in outer shell, 17 electrons in total
Cl^-	$1s^22s^22p^63s^23p^6$	1 electron gained, 8 electrons in outer shell, 18 electrons in total
Na	$1s^22s^22p^63s^1$	3 energy levels, 1 electron in outer shell, 11 electrons in total
Na^+	$1s^22s^22p^6$	1 electron lost, 8 electrons in outer shell, 10 electrons in total

First ionisation energy

First ionisation energy is the energy required to remove **one mole of electrons** from **one mole of gaseous atoms** to form one mole of gaseous **1+ ions**. It is measured in kJ mol^{-1}.

E.g. $Na(g) \rightarrow Na^+(g) + e^-$

$O (g) \rightarrow O^+(g) + e^-$

The state symbol (g) must be shown in the equation as everything has to be a gas. The ion formed must have the charge 1+.

Factors affecting ionisation energy:

- **Nuclear charge**: more protons in the nucleus means a greater nuclear charge and stronger attraction to outer shell electrons so first ionisation energy is greater.
- **Atomic radius**: a larger atomic radius means weaker attraction between the positive nucleus and negative electrons so first ionisation energy is lower.
- **Electron shielding**: more electron shells means more electron shielding so there is weaker attraction between outer shell electrons and the nucleus so first ionisation energy is lower.

The first ionisation energy of elements **increases across a period**. This is because although electron shielding remains the same, **nuclear charge increases** which draws the outer shell electrons inwards causing a slight **decrease in atomic radius**. This means more energy is required to remove an outer shell electron as it is attracted more strongly to the nucleus.

Exceptions in period 2: (These exceptions are present in period 3 for the same reasons)

- Between **groups 2 and 3**, electrons start to be **added to a 2p orbital** rather than 2s. 2p has a slightly higher energy level than the 2s orbital so this electron is found **slightly further from the nucleus**, meaning it can be removed more easily (less energy required).

- Between **groups 5 and 6**, electrons start to **pair in the 2p orbitals**. The paired electrons are both negatively charged so **repel** each other. This allows an electron to be removed with less energy than expected.

The first ionisation energy **decreases down the group** because although nuclear charge increases, the **atomic radius and electron shielding also increase** meaning less energy is required to remove an outer shell electron.

Successive ionisation energies

Successive ionisation energies involve removing one mole of electrons from **one mole of gaseous ions**. For example, the third ionisation energy of potassium would involve removing one electron from $K^{2+}(g)$ to form $K^{3+}(g)$.

The first 3 ionisation energies of aluminium are shown below:

$Al(g) \rightarrow Al^{+}(g) + e^{-}$ 1st ionisation energy = 577 kJ mol^{-1}

$Al^{+}(g) \rightarrow Al^{2+}(g) + e^{-}$ 2nd ionisation energy = 1820 kJ mol^{-1}

$Al^{2+}(g) \rightarrow Al^{3+}(g) + e^{-}$ 3rd ionisation energy = 2740 kJ mol^{-1}

Successive ionisation energies increase because atomic radius decreases and there is greater attraction between outer shell electrons and the nucleus.

A large jump between successive ionisation energies indicates which group an element is in. The successive ionisation energies for an element are shown below:

1st	2nd	3rd	4th	5th
801	2427	3660	25026	32827

The large jump between the third and fourth ionisation energy shows there are 3 electrons that are relatively easy to remove then a fourth one which requires a lot more energy to remove. This shows that there are 3 electrons in an outer shell. This means the element is in group 3. There are 5 ionisation energies for this element so it has 5 electrons. An element with 5 electrons in group 3 is boron.

Electron Affinity **(A level only)**

First electron affinity is the energy released when **one mole of gaseous atoms** each **gain an electron** to form one mole of **1- ions**. It is measured in kJ mol^{-1} and always has a negative sign to show energy is released.

E.g. $Mg(g) + e^{-} \rightarrow Mg^{-}(g)$

$Cl(g) + e^{-} \rightarrow Cl^{-}(g)$

Electron affinity **decreases down the group** because although nuclear charge increases, **electron shielding and atomic radius increase** so there is less attraction between the nucleus and an incoming electron. This means less energy is released as you go down the group. Fluorine and oxygen are exceptions to this rule as they have lower electron affinities than expected. This is because they are relatively small so are already crowded with electrons which repel an incoming electron.

Second electron affinity is the energy required to add one electron to each ion in one mole of gaseous 1- ions to form one mole of gaseous 2- ions. This requires energy because the negative ion repels the incoming electron.