CAIE Chemistry A-level

1: Atomic Structure

Notes

Particles in the Atom and Atomic Radius

Protons, Neutrons and Electrons

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**.

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.

Atomic and Mass Numbers

The **atomic number** is equal to the **number of protons** in an atom. Another name for this is the **proton number**. 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. Another name for the mass number is the **nucleon number**.

Atomic Radius

Across a **period** the **atomic radius decreases** due to an increase in protons in the nucleus, therefore an increase in nuclear charge which will **attract the outer electrons** more strongly. This pulls them closer to the nucleus and reduces the atomic radius.

Down a **group** the **nuclear charge increases** however the energy level being occupied also increases so the outer electrons are further from the nucleus. Higher energy levels are also more shielded from the nucleus by the inner shell electrons.

Ionic Radius

Across a period, the ionic radius **decreases for** positive ions then **increases for** negative ions.

The **positive** ions have a decreasing ionic radius because, although the ions have the same electron configuration, the number of protons in the nucleus increases so **nuclear attraction increases.**

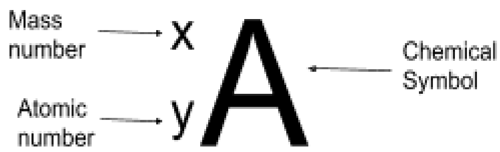
The **negative** ions have an increasing ionic radius because the ions have gained electrons meaning there are now **more electrons than protons**. As a result, the nuclear attraction to the electrons is weaker so they are not pulled in as strongly.

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:



Isotopes of the same element have the **same chemical properties** as they have the same electron configuration. However, isotopes differ in physical properties as they have different mass numbers.

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.

There are different types of orbitals:

- 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 the standard electron configuration. In the following two cases, 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.

Examples:

The **full electron configuration** of Fe: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$

The **shorthand electron configuration** of Fe: [Ar] $3d^6\ 4s^2$

The **electrons in boxes notation** of Fe:

[Ar]	↑↓	↑	↑	↑	↑		↑↓

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

Free Radical

An single unpaired electron is called a free radical, they are denoted by having a dot next to the chemical symbol, e.g. $Cl^\bullet$

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}.

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

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:

- 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.

Note, these exceptions are present in Period 3 for the same reasons.

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.