Topic 2: Atomic Structure

▼ 2.1 The Nuclear Atom

Rutherford's Experiment

- Rutherford's experiment showed that atoms have a nucleus
- It involved shooting alpha particles by a radioactive source through a sheet of gold foil, all surrounded by a florescent screen.
- As some particles were deflected back and did not go through the gold foil, it was concluded that
    - The atom was mostly empty space
    - An atom had a nucleus (proven by the deflection of particles)

Sub-Atomic Structure

- Atoms are made up of protons (with a positive charge), neutrons (with no charge) and electrons (with a negative charge)

Subatomic particle	Relative Mass	Relative Charge
Proton	1	+1
Neutron	1	0
Electron	Assumed to be 0	01

> 📜 Definitions
> Mass Number (A): Sum of number of nucleons (protons & neutrons)
> Atomic number (Z): Number of protons in the nucleus
> Isotope: Atoms of the same element with same number of protons but different number of neutrons

Nuclear notation

- Nuclear notation has a mass number, atomic number, and element symbol to represent an isotope
    - Atomic number = number of protons
    - Mass number = sum of number of protons and neutrons
    - Number of electrons = atomic number - charge

Properties of isotopes

- Chemical properties of isotopes remain the same due to an equal number of electrons (as chemical properties depend on valence electrons)
- Physical properties of isotopes (e.g. density, mass, melting, etc…) will change due to a change in the number of neutrons

Uses of radioisotopes

- Carbon-14
    - Treats cancerous cells
    - Is used in radiocarbon dating, a process to estimate organisms' age.
- Cobalt-60
    - Used in the treatment of cancer
    - Used to stop the immune responses to transplanted organs

    - Used in levelling devices
    - Used to sterilise (remove bacteria) foods and spices
- Iodine-131/Iodine-125
    - It releases gamma and beta radiation and is taken up the thyroid gland
    - Used to treat thyroid cancer and detect if the thyroid is functioning properly
    - Iodine-125 emits gamma rays and can be used to treat prostate cancer and brain tumour

Mass Spectrometry

- A mass spectrometer is used to measure individual masses of atoms. It works by separating isotopes and determining the mass of each isotopes.
- Isotopes are separated in four stages:
    1. The sample gets heated and vaporised (turns into a gas) and travels into a evacuated tube
    2. The particles will separate after passing through the evacuated tube
    3. The atoms get flooded with a stream of high energy electrons in order to knock the electrons off of the particle. This will give the particle a +1 charge.
    4. The positively charged ion will then quickly travel through the tube due to the attraction of negatively charged plates and go through slits that control its direction and velocity (speed) of its motion.
    5. The ions are then pass through a strong magnetic field, which deflects into a curved path
    6. Lastly, ions will get detected electronically to measure its location and number of particles
- The direction of the deflection of the ion depends on:
    - Absolute mass of the ion
    - The charge of the ion
    - The strength of the magnetic field

- Speed of the ions

- Information is shown in a mass spectrum by:
  - The number of peaks representing the number of isotopes
  - The position of each peak on the x-axis showing the isotopic mass
  - The height of the peaks shows the abundance (amount) of the isotopes

Calculating atomic mass

📜 A_r = (($relative\ isotopic\ mass_1$ x %$abundance_1$) + ($relative\ isotopic\ mass_2$ x %$abundance_2$) +...) $/100$

▼ 2.2 Electron configuration

Bohr's Model

- Electrons orbit the nucleus in rings at fixed energy levels. The farther away the electron, the higher the energy level
- Orbits further away from the nucleus has higher energy levels
- Electrons occurs in a ground state as they can only occur in one level energy or another, never in between
- Electrons go through an excited state when they absorb more energy to travel up to a higher energy level or travel back down to a lower energy level when emitting (releasing) energy

Emission Spectrums

- Emission spectra are made through the emission of photons from atoms as excited electron travels back to a lower energy level
- The transition between each electron travelling from a higher energy level to a lower energy level connects to a radiation with a frequency or wavelength

- Emission spectrums occurs when excited atoms travel down to a lower energy level and only containing specific colours of visible light
- Other types of spectrums:
    - Line spectrum: contains sharp and distinct colours
    - Continuous spectrum: contains all colours of visible light
- The energy of lines on an emission spectrum of hydrogen indicates the differences of energy between the energy levels
- Elements can be indicated through its emission spectrum (e.g. helium on a emission spectrum is different from hydrogen due to the difference in energy level)
- The Balmer series are from a series of spectral emission lines of a hydrogen atom due to electrons traveling from higher energy levels down to where n=2
- Ultraviolet light occurs when electrons moves down to the first energy level (Lyman Series)
- Infrared occurs when electrons moves down to the third energy levels (Paschen Series)

Electromagnetic Spectrum

- All electromagnetic waves travels at the same speed but have different wavelengths
    - Wavelength: distance between two adjacent crests
- Waves with smaller wavelengths have higher frequency; they have more energy
- Waves with longer wavelengths have lower frequency; they have less energy
- An increase in energy also increases the frequency
    - Waves with higher energy tends to cause more harm than lower energy
- Electromagnetic spectrum shows all frequencies of electromagnetic radiation
- The order from longest to shortest wavelength is: radio wave, microwave, infrared, visible, ultraviolet, x-ray, and gamma rays

Orbitals

- Orbital: A region of space in which the probability of finding an electron is greater than or equal to 95%
- However, it is impossible to locate an electron from its exact position (Heisenberg Uncertainty Principle)
- Each orbital can contain a maximum of two electrons
- The electrons within an orbital will spin in opposite directions
- Orbitals are represented as boxes with up and down arrows to depict electrons and their direction of spin (as shown with the direction of the arrow)
- The arrangement of electrons within orbitals are also determined by the following rules
    - Paull's Exclusion Principle: A maximum of two electrons (that spin in opposite directions) can occupy one orbital
    - Aufbau Principle: The electrons get placed in the orbitals with the lowest energy first
    - Hund's Third Rule: Same sub-level orbitals are first filled with one electron. Only after each sub-level has one electron can an orbital accumulate two electrons. When there are more than one sub-level orbitals available, the electron joins the orbital with an electron spinning in an opposite direction
- Orbitals are categorised into 4 main types based on their shapes
    - s-orbitals: Resemble spherical shape
    - p-orbitals: Resemble dumbbell shape
    - d-orbitals: Resemble four-leaf clover shape
    - f-orbitals: No discrete shape

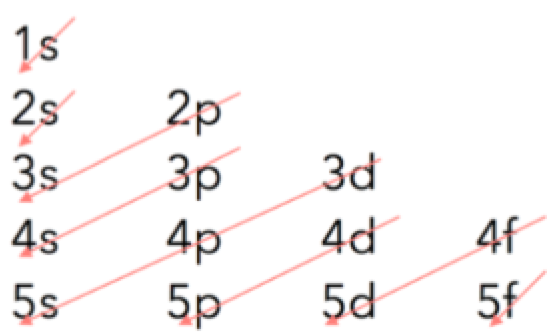
Sub-levels

- Each main energy level is divided into sub-levels

- Sub-levels are groups of a fixed number of orbitals in which electrons will be likely to be found
- The sub-levels can be categorised into s, p, d, and f based on the orbitals they contain. The orbitals can contain a maximum of 2 electrons
    - s sub-level: 1 orbital, 2 electrons
    - p sub-level: 3 orbitals: 6 electrons
    - d sub-level: 5 orbitals, 10 electrons
    - f sub-level: 7 orbitals, 14 electrons
- You can calculate the maximum amount of electrons that each main energy level (n) can hold through the expression $2n^2$

Writing electron configurations/arrangements

- The following steps help you write an electron configuration
    1. Determine the total number of electrons in the atom
    2. Fill electrons in from the lower sub-levels to the higher ones. Follow the order shown on the diagram below (eg. 1s, 2s, 2p, 3s, 3p, 4s, 3d…)



    3. Note that an S-orbital can hold 2 electrons, a p-orbital can hold 6, a d-orbital can hold 10 and a f-orbital can hold 14

💡 Example:
To write the electronic configuration of aluminum (13 electrons), here are the following steps

1. The number of electrons is 13
2. $1s^2 + 2s^2 + 2p^6 + 3s^2 + 3p^1$ (Fill up the sub-levels from the lowest to the highest)
3. Therefore, the electronic configuration of aluminum is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

As you can see, the electronic configuration of atoms can be lengthy. However, we can further simplify the electronic configuration and make it condensed through the following additional steps

1. Find the the nearest and smallest noble gas and write it in square brackets
2. Replace the electronic configuration of the noble gas with the square bracketed symbol, and continue the electronic configuration following the square bracketed symbol

💡 Example: To simplify $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$, the electronic configuration of aluminum…

1. The noble gas that is the smallest and nearest to aluminum is neon which has an electronic configuration of $1s^2\ 2s^2\ 2p^6$
2. Replace $1s^2\ 2s^2\ 2p^6$ with [Ne] to form the condensed electronic configuration
[Ne] $3s^2\ 3p^1$

The electronic configurations of the transition elements copper and chromium do not follow the expected patterns

- Electronic configuration of chromium: [Ar] $3d^5$ $4s^1$
- Electronic configuration of copper: [Ar] $3d^{10}$ $4s^1$

12.1 Electrons in atoms

📜 Definitions

First ionization energy - the minimum amount of energy required to remove one mole of electrons from one mole to gaseous atoms. The formula for first ionization energy is $X(g) \rightarrow X^+(g) + e$

- In an emission spectrum, the first ionization energy corresponds with the limit of convergence at higher frequency

First Ionization Energy Factors

Factors that influence ionization energy:

- Size of Nuclear Charge
    - The nuclear charge is the charge of the nucleus. It will increase if the atomic number (number of protons) increases
    - The attraction of electrostatic force between the nucleus and all electrons will be greater if the overall positive nuclear charge is larger
    - Removing electrons and overcoming attractive force requires a large amount of energy
    - Therefore, **first ionization energy increases across each period, as more energy is required to separate the electron from the atom**
- Distance of Outer Electrons from Nucleus
    - As distance between positive and negative charge increases, the attraction of electrostatic force between them will decreases, allowing ionization

energy to be lower

  - This is because electrons in shells farther away from the nucleus are more weakly attracted compared to electrons closer to the nucleus
  - Therefore, **Ionization energies tend to decrease down a group of the periodic table**
- Shielding Effect
  - All electrons repel each other because they are negative
  - Therefore, the electrons in the inner shells will repel full outer shell electrons
  - The shielding effect is when the electrons in full inner shell prevents the full nuclear charge from affecting the outer shell electrons
  - As shielding effect increases, the electrostatic attractive force between the nucleus and outer electrons decreases
  - **Therefore, ionization energy is lower as the number of full electron shells between the outer electron and nucleus increases due to the increased shielding effect (and resulting lessened electrostatic forces)**

First Ionization Energy Trends

- In general, first ionization energies increase from left to right across a period (more energy is required to remove electrons from left to right on a period)
- Increase of nuclear charge across a period increases the attraction between outer electrons and the nucleus, meaning more energy is required to remove a valence electron
- First ionization energy may decrease within the period
  - For example, Beryllium's electronic structure is $1s^2 2s^2$, and Boron's electronic structure is $1s^2 2s^2 2p^1$
- First ionization energy will decrease due to the orbital that the electron is removed from
  - In Beryllium, the valence electron that is being removed is from the s-orbital
  - In Boron, the valence electron that is being removed is from a p-orbital

- The p-orbital has a higher energy level. Despite the rise in attraction from the increased nuclear charge, due to the energy of the electron in the p-orbital, the overall first ionization energy is still decreased.

Successive Ionization Energies

- The second and third energies are described as:
  - $X^{+}{}_{(g)} \rightarrow X^{2+}{}_{(g)} + e^{-}$ (An ion with a +1 charge further loses one electron to gain a +2 charge)
  - $X^{2+}{}_{(g)} \rightarrow X^{2+}{}_{(g)} + e^{-}$ (An ion with a +2 charge further loses one electron to gain a +3 charge)
- Successive ionization energies increase for all atoms. This means that with each electron lost, there would need to be more ionization energy to lose the next electron
- This is because when an electron is removed, the remaining electrons will be attracted with a stronger nuclear charge.