Chapter 7: Electronic Configuration (HL)

- If an electric current is passed through a discharge tube filled with hydrogen gas a line spectrum of specific colours is observed.
- This is called a **line emission spectrum.** Spectra are studied using a machine called a **spectrometer.**
- When this is repeated with different elements, different line spectra are obtained.
- **Each element has its own unique line spectrum**.

How do line emission spectra arise?

Bohr's theory of the atom was based on the atomic emission spectrum of hydrogen.
It is as follows:

The lines appearing on the hydrogen spectrum are due to **electron transitions between energy levels**.

An energy level is the fixed amount of energy of an electron in an atom.

Atoms normally exist in the **ground state**, as this is the most stable.

The ground state of an atom is one in which the electrons occupy the lowest available energy levels.

The lowest energy level is the one closest to the nucleus and is called n=1, the next is n=2, etc. Electrons in these energy levels have fixed, 'quantised', amounts of energy.

Quantised energy means that the electrons can have only certain specific energy values. Values between the quantised values are not allowed.

When an atom receives the right amount of energy (e.g. when it is **heated** or when electricity is passed through it), **energy is absorbed**. If the **energy absorbed by an electron is equal to the difference in energy between energy levels**, electrons can use this energy to **jump from lower energy levels to higher energy levels**. This is called an **electron transition**, and the electrons are now **excited.**

The excited state of an atom is one in which the electrons occupy higher energy levels than those available in the ground state.

Electrons in the excited state are **unstable** and can't remain in the higher energy level – i.e. they fall back down to their original lower energy level.

As they fall back down, electrons **release energy in the form of a photon of light**. The energy released is the **difference in energy between the higher energy level and the lower energy level**. Since the electron is falling back down to a fixed energy level, only fixed amounts of light can be given off – **light of a definite frequency**. This specific frequency of light **corresponds to a particular colour**, which is the colour that appears on the line emission spectrum.

Photons

Photons are often described as 'packets' of energy. Photons exhibit wave–particle duality (can act as both a particle and a wave) and have a wavelength and a frequency.

Photon energy is the energy of a single photon. It is proportional to the frequency of a photon.

The frequency of a photon is defined as how many wavelengths a photon propagates per second, i.e. how many waves pass a certain point per second.

The frequency of light released can be calculated using the formula $E_m - E_n = hf$
where
E_m = the energy of the higher energy level (that the electron jumps to)
E_n = the energy of the lower energy level (that the electron falls back down to)
h = Planck's constant
f = frequency of the photon (frequency of emitted light)

• The higher the photon's frequency = the higher the energy = the shorter the wavelength.
• The lower the photon's frequency = the lower the energy = the longer the wavelength.
• **The frequency and wavelength of a photon determine the colour observed in the line emission spectrum.**

Hydrogen emission spectrum

The visible lines on the hydrogen emission spectrum occur as a result of the following transitions:

- The **violet** line occurs due to electrons falling from **n=6 to n=2.**
- The **blue-violet** line occurs due to electrons falling from **n=5 to n=2.**
- The **blue-green** line occurs due to electrons falling from **n=4 to n=2.**
- The **red** line occurs due to electrons falling from **n=3 to n=2.**

The lines occurring in the **visible region** of the hydrogen spectrum are part of the **Balmer series** (an electron transitions **from a higher energy level down to n=2**).

The **Lyman series** contains lines in the **UV region** of the spectrum, and are caused by **electrons falling to n=1**.

The **Paschen series** contains lines in the **IR portion** of the hydrogen spectrum that are caused by electrons falling from higher energy levels **to n=3**.

Bohr's theory worked well to explain the line emission spectrum of hydrogen. However, Bohr's theory couldn't satisfactorily explain emission spectra of other elements (where atoms contain multiple electrons, the effect of the electrons on each other has to be taken into account, which Bohr didn't do).

Bohr's theory also had other limitations: as we know from Chapter 6, it didn't take into account the wave nature of the electron or sublevels (since Bohr wasn't aware they existed), and his theory contradicted Heisenberg's uncertainty principle.

Experiment 2: Identifying elements using flame tests

Part 1 Flame tests

- Wooden splints are soaked overnight in water.
- A small amount of salt is crushed using a pestle and mortar.
- The wooden splint is dipped into the crushed salt and then held in the blue flame of a Bunsen burner.
- The colour produced is noted.

- This is repeated for all salts using separate pestle and mortars and separate wooden splints.

Element	Colour flame
Potassium	Lilac
Sodium	Yellow-orange
Copper	Blue-green
Lithium	Crimson
Barium	Yellow-green
Strontium	Red (scarlet)

Different colour flames produced by different elements are the results of electron transitions. As they are heated, electrons absorb energy equal to the difference between two energy levels, and use it to jump from lower energy levels to higher energy levels. As electrons fall back down to the lower energy levels, photons of light are emitted. This light corresponds to the observed colour flame.

Part 2 Line-emission spectra
You must be able to identify three elements from their line-emission spectra – sodium, strontium and copper.

Sodium
- **Characteristic bright yellow doublet** (sodium D-lines)
- Two very closely spaced lines at wavelengths of approximately 589.0nm and 589.6nm

Strontium
- Dominated by strong red lines, particularly in the 640–670nm range.
- (Also displays lines in the blue region & some orange)

Copper
- Distinct blue-green lines, notably around 510nm and 521nm.

Sublevels

As technology developed, scientists studying line emission spectra found **that what had originally seemed to be one line was actually a number of lines very close together**.

Scientists therefore reasoned that energy levels must be subdivided further – into sublevels – and that these lines must have been caused by electrons transitioning between sublevels.

A sublevel is a subdivision of a main energy level.

Below is a breakdown of energy levels and their sublevels:
n=1 1s
n=2 2s 2p
n=3 3s 3p 3d
n=4 4s 4p 4d 4f

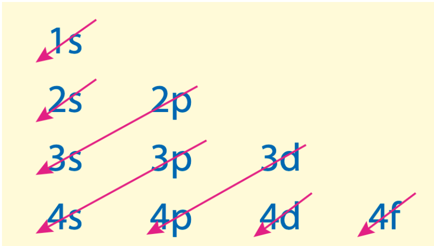
Each type of sublevel can hold a specific number of electrons. **An s sublevel can hold two electrons, a p sublevel can hold six electrons, while a d sublevel can hold ten electrons**.

Writing electronic configurations including sublevels

To write electronic configurations with sublevels, you must know three things:
1. The number of electrons in an atom of that element
2. The order in which sublevels are filled
3. The number of electrons each sublevel can hold

The Aufbau principle states that when building up the electron configuration of an atom in its ground state, the electrons occupy the lowest available energy levels.

The Aufbau (the German for 'building up') principle tells us that sublevels are filled in order of increasing energy: 1s 2s 2p 3s 3p 4s 3d 4p 4d 4f. (Note the placement of 4s before 3d).



Element	Atomic number	Electron configuration
Hydrogen (H)	1	$1s^1$
Helium (He)	2	$1s^2$
Lithium (Li)	3	$1s^2 2s^1$
Beryllium (Be)	4	$1s^2 2s^2$
Boron (B)	5	$1s^2 2s^2 2p^1$
Carbon (C)	6	$1s^2 2s^2 2p^2$
Nitrogen (N)	7	$1s^2 2s^2 2p^3$
Oxygen (O)	8	$1s^2 2s^2 2p^4$
Fluorine (F)	9	$1s^2 2s^2 2p^5$
Neon (Ne)	10	$1s^2 2s^2 2p^6$
Sodium (Na)	11	$1s^2 2s^2 2p^6 3s^1$
Magnesium (Mg)	12	$1s^2 2s^2 2p^6 3s^2$
Aluminium (Al)	13	$1s^2 2s^2 2p^6 3s^2 3p^1$
Silicon (Si)	14	$1s^2 2s^2 2p^6 3s^2 3p^2$
Phosphorus (P)	15	$1s^2 2s^2 2p^6 3s^2 3p^3$
Sulfur (S)	16	$1s^2 2s^2 2p^6 3s^2 3p^4$
Chlorine (Cl)	17	$1s^2 2s^2 2p^6 3s^2 3p^5$
Argon (Ar)	18	$1s^2 2s^2 2p^6 3s^2 3p^6$
Potassium (K)	19	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Calcium (Ca)	20	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Scandium (Sc)	21	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
Titanium (Ti)	22	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
Vanadium (V)	23	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$
Chromium (Cr)	24	$1s^2 2s^2 2p^6 3s^2 3p^6$ $\mathbf{4s^1 3d^5}$
Manganese (Mn)	25	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
Iron (Fe)	26	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
Cobalt (Co)	27	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$
Nickel (Ni)	28	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$
Copper (Cu)	29	$1s^2 2s^2 2p^6 3s^2 3p^6$ $\mathbf{4s^1 3d^{10}}$

Zinc (Zn)	30	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$
Gallium (Ga)	31	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$
Germanium (Ge)	32	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$
Arsenic (As)	33	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$
Selenium (Se)	34	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$
Bromine (Br)	35	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
Krypton (Kr)	36	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

Exceptions

- Chromium has an electronic configuration of $1s^2 2s^2 2p^6 3s^2 3p^6$ **$4s^1 3d^5$,** rather than $4s^2 3d^4$.
- Copper has an electronic configuration of $1s^2 2s^2 2p^6 3s^2 3p^6$ **$4s^1 3d^{10}$** rather than $4s^2 3d^9$.
- **Half-filled or fully filled sublevels have extra stability**, so copper and chromium rearrange their electrons to achieve this.

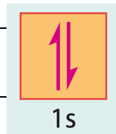
Orbitals

An orbital is a region of space around the nucleus of an atom within which there is a high probability of finding an electron.

Sublevels contain orbitals. **An orbital can hold two electrons.**

s sublevel = one orbital
p sublevel = three dumb-bell shaped orbitals - **p_x, p_y, and p_z,**
d sublevel = ten orbitals

Pauli's exclusion principle states that no more than two electrons may occupy an orbital, and they must have opposite spin.



Energy level	Sublevels	No. of orbitals in each sublevel	Names of orbitals	No. of electrons each sublevel holds
n=1	s	1	s	2
n=2	s, p	1, 3	p_x, p_y, p_z	2, 6
n=3	s, p, d	1, 3, 5	-	2, 6, 10
n=4	s, p, d, f	1, 3, 5, 7	-	2, 6, 10, 14

Hund's rule of maximum multiplicity states that when two or more orbitals of equal energy are available, the electrons occupy them singly before filling them in pairs.

Write the electronic configuration of oxygen, including sublevels.

Oxygen has eight electrons, which we have been writing as $1s^2 2s^2 2p^4$. However, to include sublevels we must now write this as $1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$.

State the number of (a) energy levels; (b) sub-levels; (c) orbitals occupied by electrons in a chlorine atom in its ground state.

To answer this question it's helpful to write out the electronic configuration of chlorine.

$1s^2\ 2s^2\ 2p_x^2\ 2p_z^2\ 2p_y^2\ 3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^1$

Now, we simply count.

(a) Energy levels = 1, 2, 3 = 3 energy levels

(b) Sublevels = 1s, 2s, 2p, 3s, 3p = 5 sublevels

(c) Orbitals = 1s 2s $2p_x$ $2p_y$ $2p_z$ 3s $3p_x$ $3p_y$ $3p_z$ = 9 orbitals

Write the electronic configuration (including sublevels) of Na^+.

$[1s^2\ 2s^2\ 2p^6]^+$

Note: If asked to include orbitals as well as sublevels, we would write: $[1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2]^+$

Write the full electronic configuration (showing energy levels, sublevels and orbitals) of Cl^-.

$Cl^- = [1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2\ 3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^2]^-$