Structure 1.3

IB CHEMISTRY SL

Structure 1.3.1

Understandings:

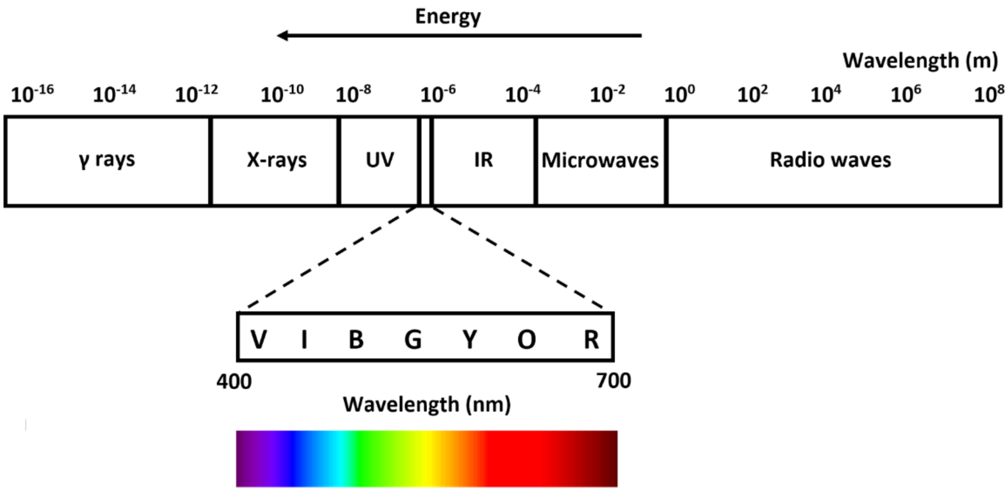
- Emission spectra are produced by atoms emitting photons when electrons in excited states return to lower energy levels.

Learning outcome(s):

- Qualitatively describe the relationship between colour, wavelength, frequency and energy across the electromagnetic spectrum.
- Distinguish between a continuous and a line spectrum.

The electromagnetic spectrum

- The electromagnetic spectrum is the range of wavelengths, or frequencies, of electromagnetic radiation.
- It extends from radio rays (low energy, long wavelength, low frequency) to gamma rays (high energy, short wavelength, high frequency).



- Higher energy corresponds to higher frequency and shorter wavelength.
- Lower energy corresponds to lower frequency and longer wavelength.

Exercises:

1. Write the following in order of increasing energy.

   UV, visible light, gamma rays, X-rays, microwaves, radio waves, infrared

2. Write the following in order of increasing wavelength.

   UV, visible light, gamma rays, X-rays, microwaves, radio waves, infrared

3. Write the following in order of increasing frequency.

   UV, visible light, gamma rays, X-rays, microwaves, radio waves, infrared

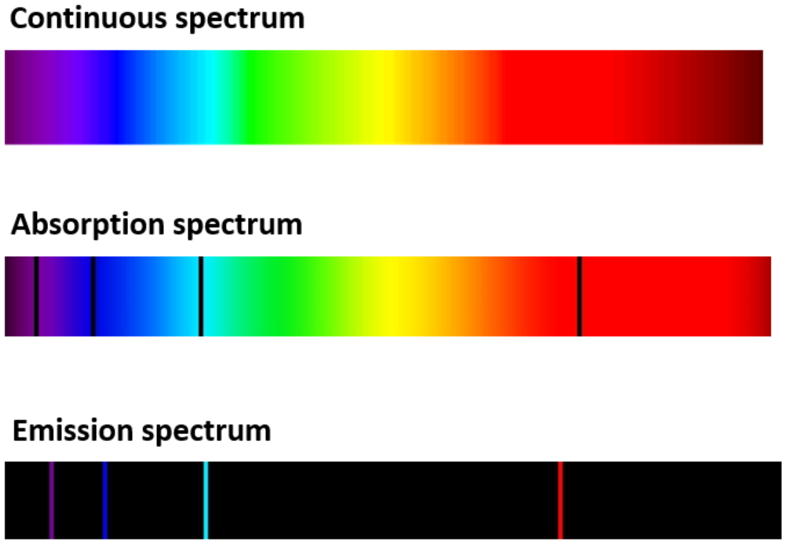
4. Write the following in order of increasing energy.

   Orange, yellow, red, violet, green, indigo, blue

5. State the relationship between energy, frequency and wavelength.

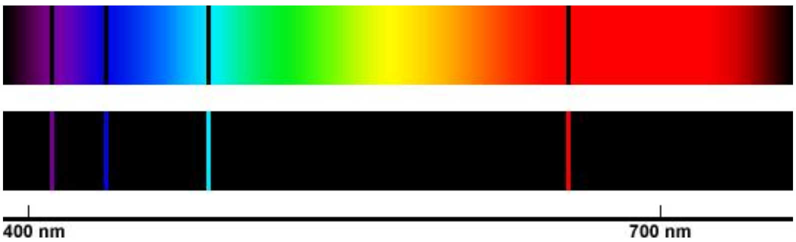
Line spectra

- The three types of line spectra are continuous, absorption, and emission spectra.



- A continuous spectrum shows all the wavelengths, or frequencies, of visible light.
- An absorption spectrum shows black lines on a coloured background.
- An emission spectrum shows coloured lines on a black background.
- Each element has unique absorption and emission spectra and they can be used to identify unknown elements.

Exercises

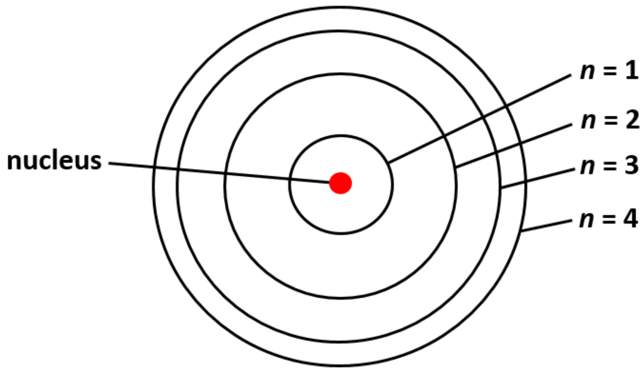


1. Classify the spectra above as absorption or emission spectra.

2. Describe the difference between the two spectra.

How are line spectra produced?

- The Bohr model of the atom has the protons and neutrons located in the nucleus and the electrons located in energy levels around the nucleus.



- Electrons can only exist within the energy levels and electrons in the same energy level have the same amount of energy.
- Electrons can transition between energy levels by either absorbing or emitting specific amounts of energy.
- The energy is in the form of small packets of energy called photons.
- If an electron absorbs an exact amount of energy, it will transition to a higher energy level (for example from n = 1 to n = 2).
- If an electron emits an exact amount of energy, it will transition to a lower energy level (for example n = 4 to n = 2).

Structure 1.3.2
Understandings:

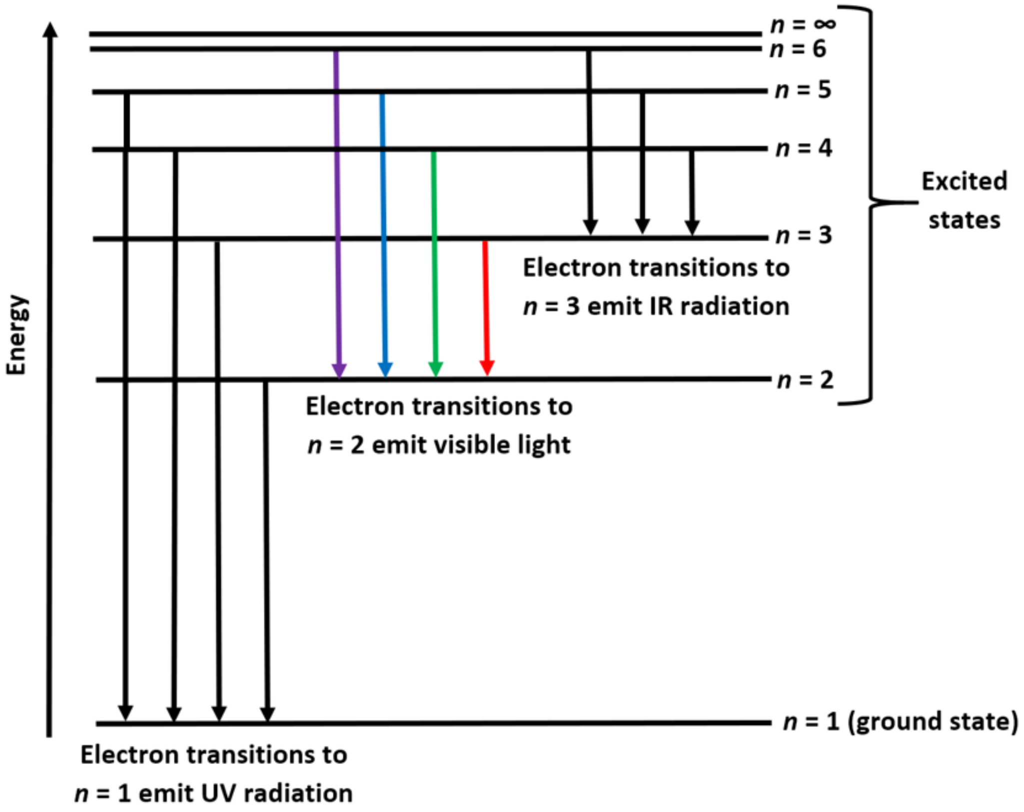
- The line emission spectrum of hydrogen provides evidence for the existence of electrons in discrete energy levels, which converge at higher energies.

Learning outcome(s):

- Describe the emission spectrum of the hydrogen atom, including the relationships between the lines and energy transitions to the first, second and third energy levels.

The hydrogen emission spectrum

- The hydrogen emission spectrum is shown below.



- Electron transitions to the first energy level (*n*=1) release the highest amount of energy and are in the UV region of the electromagnetic spectrum.
- Electron transitions to the *n*=2 energy level emit energy that corresponds to the frequency, or wavelength of visible light.
- Electron transitions to the *n*=3 energy level emit energy in the infrared region of the electromagnetic spectrum.
- The longer the arrow, the greater the amount of energy emitted (or absorbed).
- Higher energy corresponds to higher frequency and shorter wavelength.
- Lower energy corresponds to lower frequency and longer wavelength.

Exercises:

1. What is absorbed when an electron transitions from a lower energy level to a higher energy level?

2. What is emitted when an electron transitions from a higher energy level to a lower energy level?

3. Do spectral lines converge at high energy or low energy?

4. Electron transitions to n=1 emit which type of electromagnetic radiation?

5. Electron transitions to n=2 emit which type of electromagnetic radiation?

6. Electron transitions to n=3 emit which type of electromagnetic radiation?

Structure 1.3.3 and 1.3.4

Understandings:

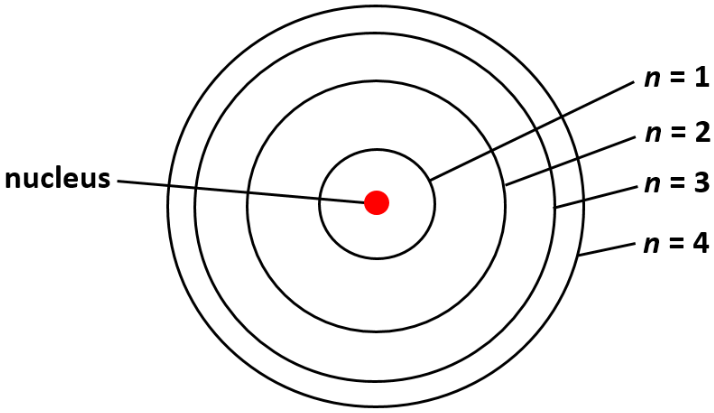
- The main energy level is given an integer number, n, and can hold a maximum of $2n^2$ electrons (1.3.3).
- A more detailed model of the atom describes the division of the main energy level into s, p, d and f sublevels of successively higher energies (1.3.4).

Learning outcome(s):

- Deduce the maximum number of electrons that can occupy each energy level (1.3.3).
- Recognize the shape and orientation of an s atomic orbital and the three p atomic orbitals (1.3.4).

Electron configurations

- The Bohr model of the atom has the electrons located in energy levels (principal energy levels) which are assigned the letter *n*.



- *n*=1 is closest to the nucleus and has the lowest energy. As the value of *n* increases, the energy also increases.
- Each main energy level can hold $2n^2$ electrons.
- The main energy levels are divided into sub-levels: s, p, d and f.
- The order in terms of energy of the sub-levels is: s < p < d < f (s is lowest and f is highest).

Energy level	sub-level	maximum number of electrons in sub-level	maximum number of electrons in level
$n = 1$	1s	2	2
$n = 2$	2s	2	8
	2p	6	
$n = 3$	3s	2	18
	3p	6	
	3d	10	
$n = 4$	4s	2	32
	4p	6	
	4d	10	
	4f	14	

Exercise:

- The *n=1* energy level can hold a maximum of ___ electrons.
- The *n=2* energy level can hold a maximum of ___ electrons.
- The *n=3* energy level can hold a maximum of ___ electrons.
- The *n=4* energy level can hold a maximum of ___ electrons.

Structure 1.3.5

Understandings:

- Each orbital has a defined energy state for a given electron configuration and chemical environment, and can hold two electrons of opposite spin.
- Sublevels contain a fixed number of orbitals, regions of space where there is a high probability of finding an electron.

Learning outcome(s):

- Apply the Aufbau principle, Hund's rule and the Pauli exclusion principle to deduce electron configurations for atoms and ions up to $Z = 36$.

Additional notes:

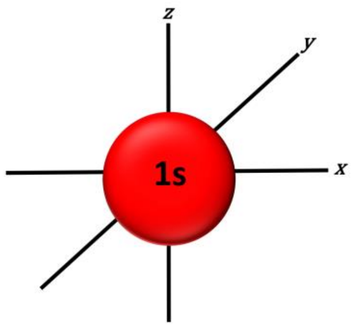
- Full electron configurations and condensed electron configurations using the noble gas core should be covered.
- Orbital diagrams, i.e. arrow-in-box diagrams, should be used to represent the filling and relative energy of orbitals.
- The electron configurations of Cr and Cu as exceptions should be covered.

Atomic orbitals

- Atomic orbitals describe the probability of finding an electron in an area of space.
- They represent the region around the nucleus where there is a 95% chance of finding an electron.

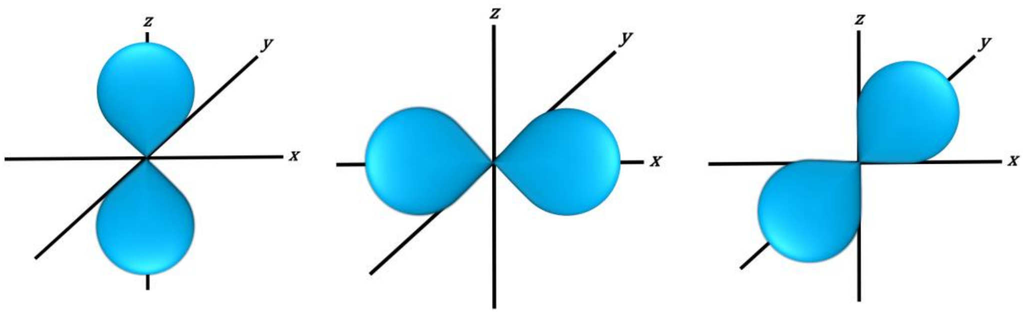
s atomic orbitals

- s orbitals are spherical in shape and can hold a maximum of two electrons.



p atomic orbitals

- A p orbital is like two identical balloons tied together at the centre (dumbbell shaped).
- The p sub-level contains three p orbitals of equal energy (degenerate orbitals) and can hold a maximum of six electrons.

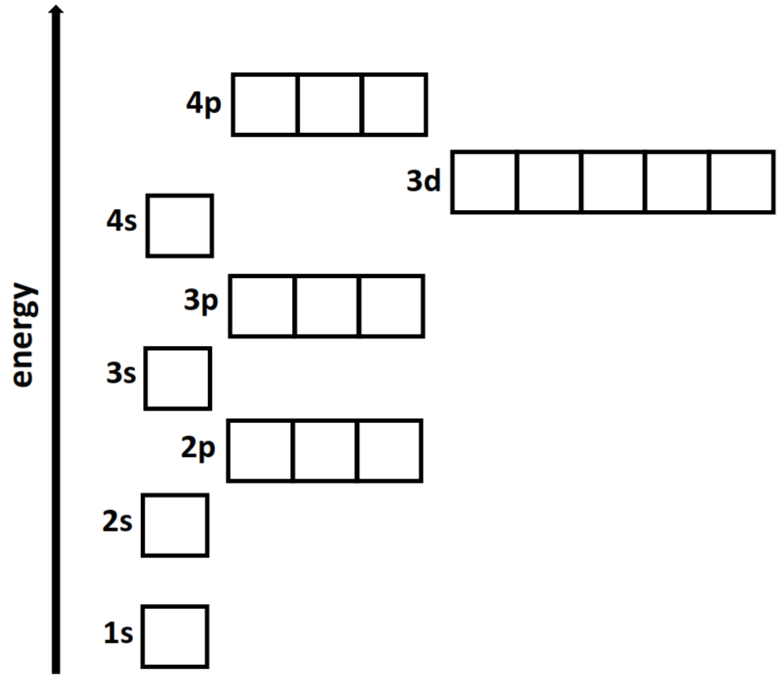


d and f atomic orbitals

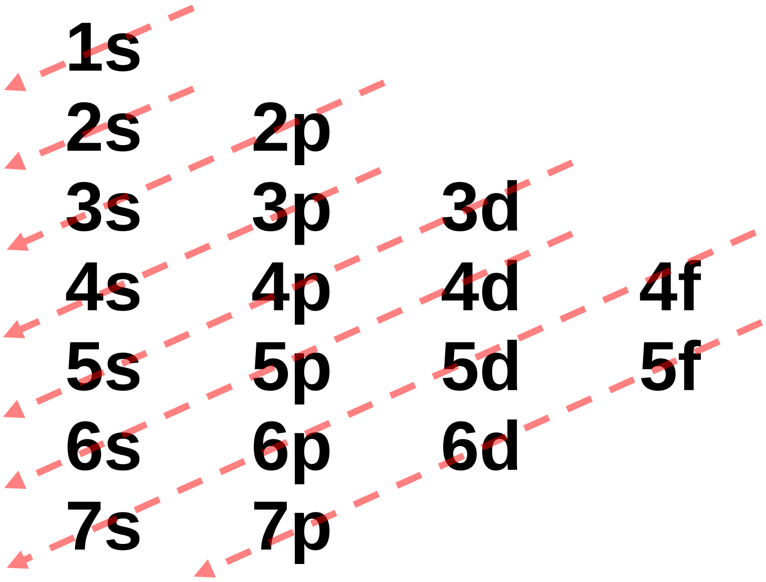
- The d sub-level contains five degenerate d orbitals and can hold a maximum of 10 electrons.
- The f sub-level contains seven degenerate f orbitals and can hold a maximum of 14 electrons.
- Students are not required to know the shapes of d and f atomic orbitals.

The Aufbau Principle

- The Aufbau Principle states that electrons are placed into orbitals of lowest energy first.
- The following diagram shows the sub-levels in order of increasing energy.
- Note the overlap between the 4s and 3d sub-levels.



- The filling of the sub-levels follows the pattern below.



Electron spin and the Pauli Exclusion Principle

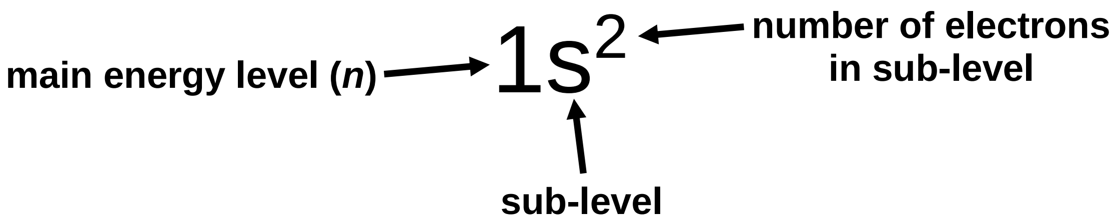
- The Pauli Exclusion Principle states that no two electrons in the same orbital can have the same quantum number.
- This means that no more than two electrons can occupy an orbital and they must spin in opposite directions.
- Electrons and their spins are represented by single-headed arrows (↿ or ⇂).

Hund's rule

- Hund's rule states that if more than one degenerate orbital in a sub-level is available, electrons occupy separate orbitals with parallel spins.
- Always fill orbitals of equal energy with one electron first and then add the second electron once each orbital has one electron in it.

Writing electron configurations

- Electron configurations show how electrons are arranged in sub-levels.
- The first number shows the main energy level (or principal quantum number).
- The letter shows the sub-level (s, p, d or f).
- The number in superscript shows the number of electrons in the sub-level.



Condensed electron configurations

Notation	Electron configuration (core electrons)
[He]	$1s^2$
[Ne]	$1s^2\ 2s^2\ 2p^6$
[Ar]	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
[Kr]	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

Example: the full and condensed electron configurations of rubidium (Rb) are shown below.

- Rb $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^1$
- Rb [Kr] $5s^1$

Exercises:

1. Write full electron configurations for the following atoms:

1) He	11) Ar
2) Li	12) Ca
3) B	13) Ti
4) C	14) Mn
5) O	15) Ni
6) Ne	16) Zn
7) Na	17) Ge
8) Al	18) Se
9) P	19) Br
10) Cl	20) Kr

2. Write condensed electron configurations for the following atoms.

1) Li
2) Mg
3) S
4) Ca
5) Ga

Electron configurations of ions

- Note that First row d-block elements (Sc to Zn) lose their 4s electrons first when they form ions.

  Write the condensed electron configuration for the Ni^{2+} ion.

  Write the condensed electron configuration for the Mn^{2+}ion.

Exercise: write condensed electron configurations for the following ions:

1) Na^{+}
2) S^{2-}
3) Ca^{2+}
4) Cr^{3+}
5) Cu^{+}

Exceptions to the Aufbau principle: copper (Cu) and chromium (Cr)

Chromium Z=24

- The abbreviated electron configuration for the Cr atom is:

Copper Z=29

- The abbreviated electron configuration for the Cu atom is:

Orbital diagrams – arrows in boxes

- Boxes can be used to represent the atomic orbitals with single headed arrows used to represent the spinning electrons.
- Recall that electrons fill orbitals according to Hund's rule and the Pauli exclusion principle; an orbital can hold a maximum of two electrons which must have opposite spins, ↿ or ⇂, and degenerate orbitals are filled singly before being doubly occupied.

Exercises:

1. Draw arrows in boxes (orbital diagrams) for the first 7 elements below:

H He Li Be Boron

Carbon Nitrogen Oxygen

2. Draw orbital diagrams for the following showing only the 4s and 3d sub-levels.

1. Ca
2. V
3. Mn
4. Cr^{3+}
5. Cu^{2+}