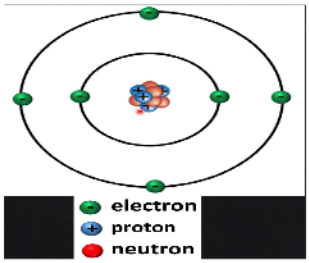
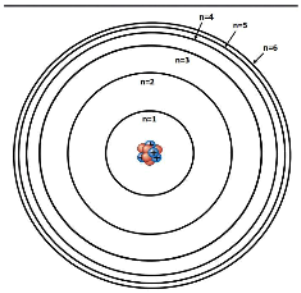
Group 2

Charlie- Subatomic particles and structures of an atom, Atomic number and mass number, Isotopes

- **Subatomic particles and structure of an atom**
  - Understandings:
    - Atoms contain a positively charged dense nucleus composed of protons and neutrons (nucleons).
    - Negatively charged electrons occupy the space outside the nucleus.
    - The main energy level or shell is given an integer number, *n*, and can hold a maximum number of electrons, $2n^2$.
    - A more detailed model of the atom describes the division of the main energy level into s, p, d and f sub-levels of successively higher energies.
  - Summaries
    - The negatively charged electrons are located in energy levels surrounding the nucleus



    - Each main energy level can be split into sub-levels



main energy level (n)	sub-level	number of electrons in sub-level	number of electrons in main energy level
1	1s	2	2
2	2s	2	8
	2p	6	
3	3s	2	18
	3p	6	
	3d	10	
4	4s	2	32
	4p	6	
	4d	10	
	4f	14	

- **Atomic Number and Mass number**
    - Applications and skills
        - Use of the nuclear symbol notation to deduce the number of protons, neutrons and electrons in atoms and ions.
    - Summary
        - Atomic Number- The number of protons in an Atom.
        - Mass Number- The number of protons plus the number of neutrons in the nucleus of an Atom.

$^{A}_{Z}X$

X - symbol of element
A - mass number
Z - atomic number

- **Isotopes**
    - Understandings
        - Definition and properties of Isotopes
    - Summary
        - Isotopes are atoms of the same element that have the same number of protons but a different number of neutrons.

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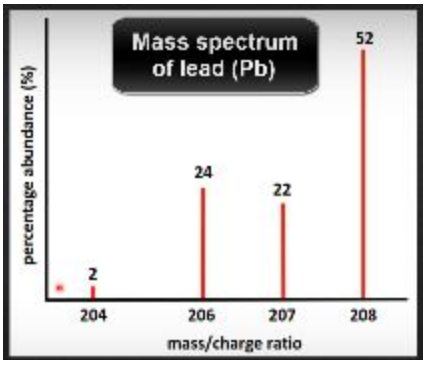
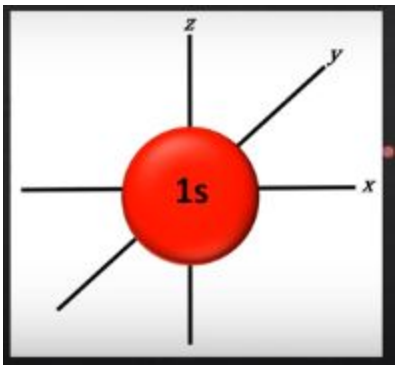
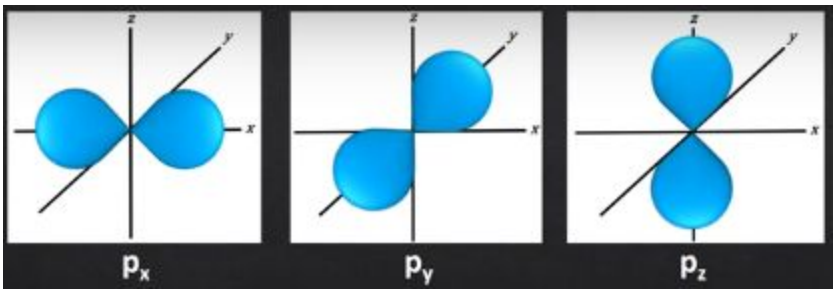
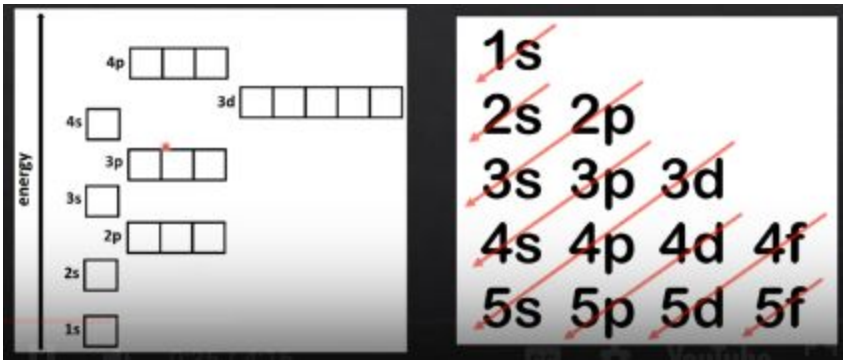
- Calculating Relative atomic mass
    - **Understandings:**
        - The mass spectrometer is used to determine the relative atomic mass of an element from its isotopic composition.
    - **Applications and skills:**
        - Calculations involving non-integer relative atomic masses and abundance of isotopes from given data, including mass spectra.
    - Summary:
        - To calculate relative atomic mass, you must multiply the percent abundance of each isotope by the corresponding mass.
        - 

        - A_r = (0.02)(204)+(0.24)(206)+(0.22)(207)+(0.52)(208) = 207.20
- Atomic Orbitals
    - **Understandings:**
        - Sub-levels contain a fixed number of orbitals, regions of space where there is a high probability of finding an electron.
    - **Applications and skills:**
        - Recognition of the shape of an s atomic orbital and the px, py and pz atomic orbitals.
    - Summary:
        - 



        - Orbitals represent a region of space where there is a high chance of finding an electron there.
        - s orbitals can hold a maximum of 2 electrons with opposite spins.
        - p orbitals have three components: p_x, p_y, and p_z
            - Each can have a maximum of two electrons with opposite spins (total of 6 electrons)
- The Aufbau Principle
    - **Understandings:**
        - Each orbital has a defined energy state for a given electronic configuration and chemical environment and can hold two electrons of opposite spin.
    - **Applications and skills:**
        - Application of the Aufbau principle, Hund's rule and the Pauli exclusion principle to write electron configurations for atoms and ions up to Z = 36.
    - Summary:
        - Electrons will fill the lowest energy orbitals first.

- Pauli Exclusion Principle: an atomic orbital can hold two electrons with opposite spins
- Hund's Rule: orbitals in a sub-level must all be singly occupied before a second electron can be added
- 

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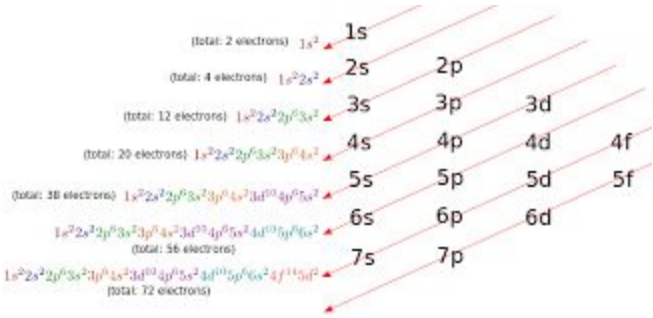
- Electron Configurations

Applications and Skills:
Application of the Aufbau principle, Hund's rule and the Pauli exclusion principle to write electron configurations for atoms and ions up to Z = 36.

Summary:

+ An atomic orbitals can hold a maximum of 2 electrons with opposite spins.
+ 4 sub-level fills before the 3rd sub-level.



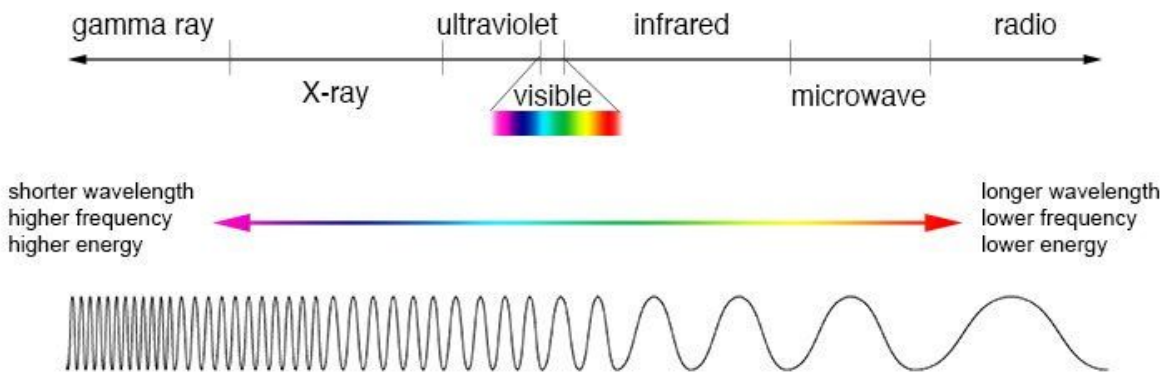
- Electromagnetic Spectrum

Applications and Skills:
Description of the relationship between colour, wavelength, frequency and energy across the electromagnetic spectrum.

Summary:

+ Speed of light= frequency* wavelength

+ High energy= High Frequency = short wavelength
+ Low energy= low frequency = long wavelength

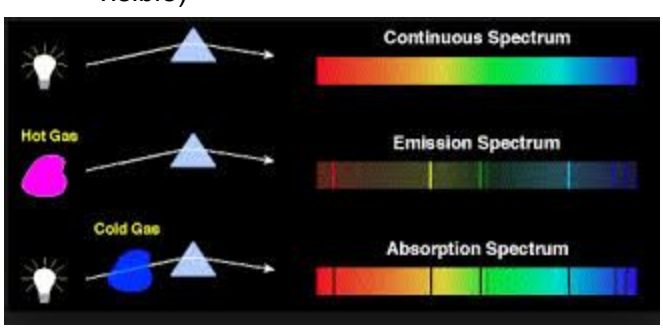


- Line Spectra

Understandings:
3 types of spectrum

Applications and Skills:
Distinction between a continuous spectrum and a line spectrum.

Summary:
+ Continuous spectrum- shows all wavelength of visible light.
+ Absorption line spectrum- black lines on coloured background (certain wavelength missing)
+ Emission line spectrum- Coloured lines on a black background (only certain wavelength visible)



- **Hydrogen Emission Spectrum**
  - **Understandings:**

    Emission spectra are produced when photons are emitted from atoms as excited electrons return to a lower energy level.

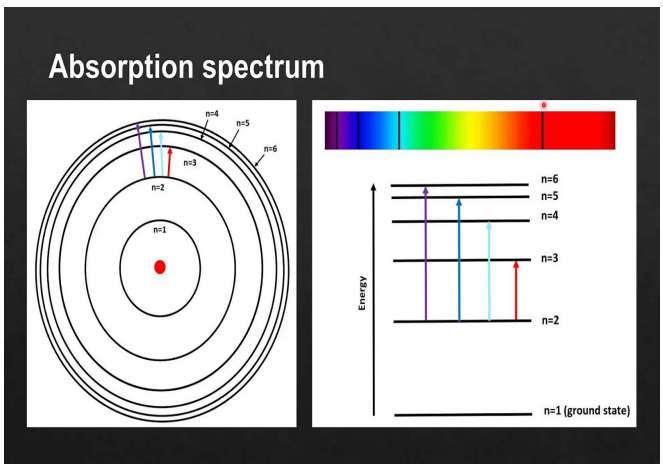
The line emission spectrum of hydrogen provides evidence for the existence of electrons in discrete energy levels, which converge at higher energies.

- **Applications and skills:**

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

- **Summary**

- Electrons transition between energy levels. When an electron transfers to higher levels, they are in an excited state
- Energy is gained or lost in the form of photons
- The amount of energy absorbed corresponds to the wavelength of a color of light
- When electrons fall to lower energy levels, the energy is released, emitting light
- Electron transitions in the emission spectrum of hydrogen emit UV radiation when electrons transition to the 1st energy level, visible light when electrons transition to the 2nd energy level, and infrared radiation when they transition to the 3rd energy level
- The spectral lines converge at high energy because there is less of a difference in energy between the levels



- **Exceptions to the Aufbau principle**
  - Understandings
    - This video covers exceptions to the Aufbau principle (Cu and Cr) as well as writing abbreviated electron configurations.
  - Summary
  - There are two exceptions to the Aufbau principle : Copper and Chromium
  - Chromium has a half filled 4s orbital
  - Copper has a half filled 4s and 3d orbital
  - You can use noble gas notation to abbreviate electron configurations by starting where the closest noble gas ends (Putting the symbol of the noble gas in brackets) , then writing the rest of the configuration
  - When writing the electron configuration of a d block ion, they lose the electrons in their 4s orbital first

Element	Should be	Actually is
Copper	$1s^2 2s^2 2p^6 3s^2 3p^6 \underline{3d^4 4s^2}$	$1s^2 2s^2 2p^6 3s^2 3p^6 \underline{3d^5 4s^1}$
Chromium	$1s^2 2s^2 2p^6 3s^2 3p^6 \underline{3d^9 4s^2}$	$1s^2 2s^2 2p^6 3s^2 3p^6 \underline{3d^{10} 4s^1}$

- **Orbital Diagrams**
    - Understandings:

Each orbital has a defined energy state for a given electronic configuration and chemical environment and can hold two electrons of opposite spin.

- Summary
- Use both the Aufbau principle and the Pauli exclusion principle to define how to notate the electrons in orbital notation.

Oxygen