Topic 2 Atomic Structure SL

2.1/2.2 Sub-atomic particles and structure of an atom

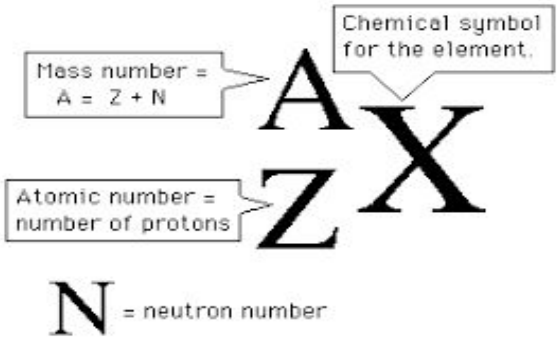
Objectives: Understandings: Atoms contain a positively charged dense nucleus composed of protons and neutrons (nucleons).
Negatively charged electrons occupy the space outside the nucleus.
The main energy level or shell is given an integer number, *n*, and can hold a maximum number of electrons, 2*n*2.
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.

Summary: Protons and neutrons are located in the nucleus of the atom. Electrons are located in energy levels surrounding the nucleus. Protons have a relative mass of 1 and a charge of +1. Neutrons have a relative mass of 1 and a charge of 0. Electrons have a relative mass of 1/2000 and a charge of -1. Most of the mass of the atom comes from the nucleus. Energy levels can be split into sub-levels. The main energy level that is closest to the nucleus has the lowest energy. Levels can be represented as “n”. The first energy level (n=1) has 1 sub level, 1s. The second energy level (n=2) has 2 sub-levels: 2s and 2p. The third energy level (n=3) as 3 sub-levels: 3s, 3p, 3d. The fourth energy level (n=4) has 4 sub-levels: 4s, 4p, 4d, 4f. Energy level can hold up to 2n^2 electrons, n representing the energy level.

2.1 Atomic number and mass number

Objectives: Applications and skills: Use of the nuclear symbol notation to deduce the number of protons, neutrons and electrons in atoms and ions.

Summary: The atomic number of an element is the number of protons in its nucleus. The mass number is the number of protons plus the number of neutrons. The charge is equal to the sum of the protons (counting as +1) and the electrons (having -1 charge). In nuclear symbol notation, the atomic number is bottom left, with the mass number above it.



2.1 Isotopes

Objectives: Definition and properties of isotopes

Summary: Isotopes are atoms of the same element that have a different number of neutrons (and different mass number).

2.1 Calculating relative atomic mass

Objectives: 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: Multiply percent abundance to the mass of isotope, add them together and divide it by 100 to find relative atomic mass

2.2 Atomic orbitals

Objectives: 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: the S and P sublevels have 1 and 3 orbitals, and can hold 2 and 6 electrons respectively. Since the location of an electron cannot be known, the shape of the orbital represents where the electrons most likely are. The single S orbital is shaped like a sphere around the nucleus, the 3 P orbitals are shaped like dumbbells in 3 different orientations of space.

2.2 The Aufbau principle

Objectives: 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: Electrons will always fill the lowest energy orbital, starting with 1s. The order is not always determined by sublevel, for example 4s is lower energy than 3d. Each orbital can only be filled with two electrons of opposite spin, but there must always be at least 1 electron in each orbital of the sublevel before the second one can be in the orbital.

2.2 Electron configurations

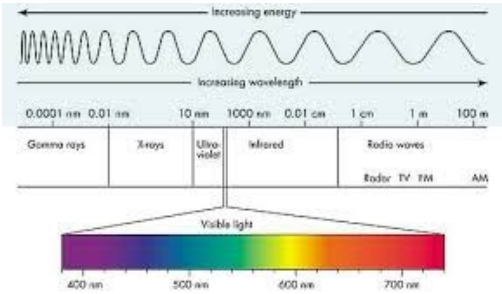
Objectives: Applications and skills: Application of the Aufbau principle, Hund's rule and the Pauli exclusion principle to write electron configurations for atoms and ions up to Z = 36.
Guidance: Orbital diagrams should be used to represent the character and relative energy of orbitals.
The electron configurations of Cr and Cu as exceptions should be covered.

Summary: the format used for electron configurations is as follows: Main energy level number, sublevel (e.g. s, p, d, f), number of electrons in sublevel. Another notation is noble gas notations, this is where the closest noble gas is used instead of the actual configuration up to that point.

2.2 Electromagnetic Spectrum

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

Summary: Electromagnetic spectrum shows all the frequencies of electromagnetic radiation. Length of wave is inversely proportional to frequency, the higher the frequency the shorter the wavelength and vise versa. Radiation having the lowest frequency to microwaves, infrared, visible light, ultraviolet, x-rays, to gamma rays. $c=\lambda v$ shows the relationship between frequency and wavelength, c is the speed of light being a constant, λ being wavelength, and v being frequency.

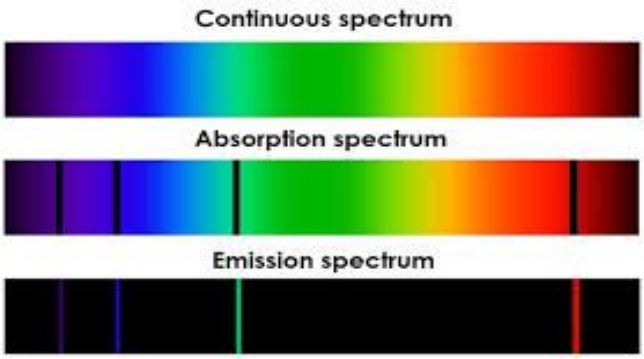


2.2 Line Spectra

Objectives: Applications and skills: Distinction between a continuous spectrum and a line spectrum.

Summary: A continuous spectrum shows all wavelengths of visible light, an absorption line spectrum shows a colored background with certain wavelengths missing as black

lines. A line spectrum (emission spectrum) is colored lines on a black background where only certain wavelengths are visible.



2.2 Hydrogen emission spectrum

Objectives: Understandings:

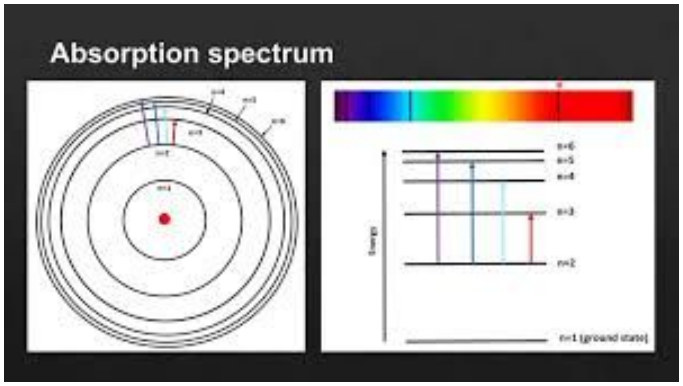
Emission spectra are produced when photons are emitted from atoms as excited electrons return to a lower energy level.

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 can transition between energy levels. Spectral lines are produced from electrons transitioning to different energy levels. If an electron absorbs energy it can transition to a higher energy level (excited state). Visible light results from energy transition down to the second energy level.



2.2 Exceptions to the Aufbau principle

Objectives: Exceptions to the Aufbau principle (Cu and Cr) as well as writing abbreviated electron configurations.

Summary: Square boxes represent orbitals, and sometimes there are some zest exceptions such as Copper where one of the 4s electrons is actually in the 3d same with Chromium.

2.2 Orbital diagrams

Objectives: Understandings:
Each orbital has a defined energy state for a given electronic configuration and chemical environment and can hold two electrons of opposite spin.
Guidance:
Orbital diagrams should be used to represent the character and relative energy of orbitals.

Summary: The aufbau principle and Hund's rule tells us how to fill electron orbitals; each orbital is represented by a box with electrons represented by arrows. Due to the aufbau principle, each sublevel must be filled before starting on the next one. Due to Hund's rule each orbital in the sublevel must have 1 electron before any can have 2.