Live Revision

Grade 11 chemistry

Unit 1

Atomic Structure and the Periodic Table

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1.1 Introduction

- The concept of the atom originated with Democritus (460–370 BC), who suggested that matter is made up of indivisible particles called "atomos".
- Aristotle disagreed with Democritus and believed matter was continuous, a view that was accepted for centuries.
- In 1803, John Dalton proposed the modern atomic theory, stating that atoms are indivisible particles that combine to form matter.

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Dalton's atomic theory

- The laws that are the basis for Dalton's atomic theory are the law of conservation of mass and the law of definite proportions.

- What is the law that Dalton formulated based on the law of conservation of mass and the law of definite proportions?

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- **Postulates of Dalton's atomic theory**
  - Elements are made of small particles called atoms.
  - Atoms are indivisible
  - Atoms can neither be created nor destroyed.
  - All atoms of the same element are identical and have the same mass and size.
  - Atoms of different elements have different masses and size.
  - Atoms combine in small whole numbers to form compounds

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- **Postulates of the modern atomic theory**
  - Atoms are the smallest particles of all elements
  - An atom is divisible. It can be subdivided into electrons, protons, and neutrons.
  - An atom is also indestructible i.e., atoms can neither be created nor destroyed during ordinary chemical reactions.
  - Atoms of the same element may not be identical in mass because of the existence of isotopes.
  - Atoms of the same elements have identical chemical properties.
  - Atoms of different elements have different chemical properties.
  - Atoms of two or more elements combine in simple whole-number ratios to form compounds.

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1.3 Early experiments to characterize the atom

Discovery of electron

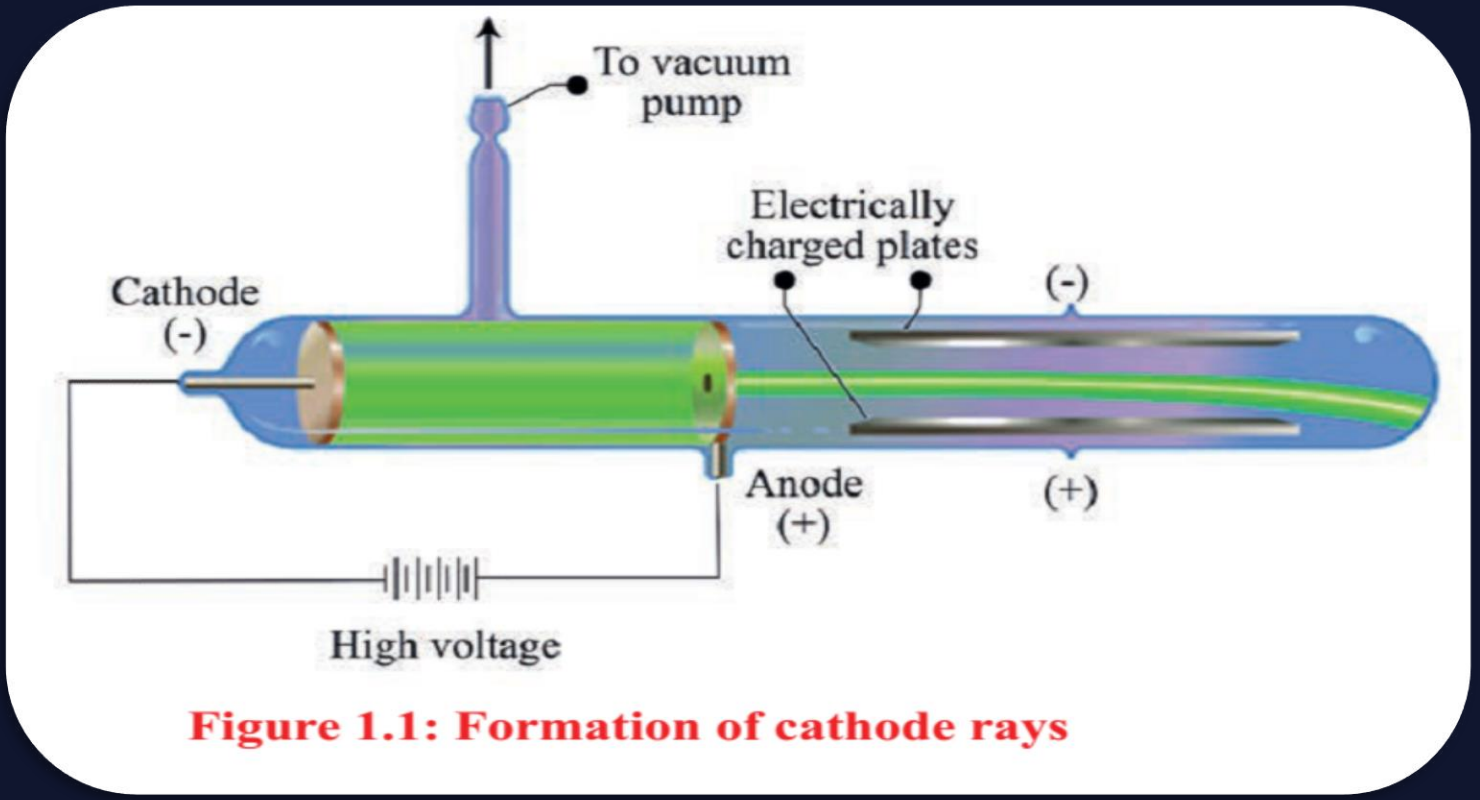


Figure 1.1: Formation of cathode rays

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Cathode Ray and Electron

- **Cathode rays** are streams of negatively charged particles emitted from the **cathode** (negative electrode) in a vacuum tube.
- The **anode** (positive electrode) attracts cathode rays, indicating the rays carry a negative charge.
- **J.J. Thomson** (1897) discovered that cathode rays are made of subatomic particles called **electrons**.
- He measured the charge-to-mass ratio (e/m) of electrons and found it to be 1.76×10^{11} C/kg(5.686×10^{-12}kg C^{-1} .
- **Robert Millikan's** oil-drop experiment (1909) determined the charge of an electron to be -1.6×10^{-19} C.
- Using both discoveries, the mass of an electron was calculated as 9.1×10^{-31} kg.
- These experiments proved that atoms are divisible and contain negatively charged particles (electrons).

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Radioactivity and Discovery of Nucleus

Radioactivity

- **Radioactivity** or **radioactive decay** is the spontaneous emission of particles and/or radiation from the unstable nuclei of certain atoms such as uranium, radium, etc.
- There are three types of radiation emitted during radioactive decay:
  - **Alpha (α) particles** – Positively charged, low penetration.
  - **Beta (β) particles** – Negatively charged, moderate penetration.
  - **Gamma (γ) rays** – Electromagnetic radiation, high penetration.
- Two are deflected by oppositely charged metal plates.

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- **1. Alpha(α) particles** Alpha(α)rays consist of positively charged particles
  - They have a mass of about four times that of a hydrogen atom
  - A charge twice the magnitude of an electron; they are identical to helium nuclei.
- **2. Beta (β) particles**
  - β particles, are electrons coming from inside the nucleus and are deflected by the negatively charged plate.
- **3. Gamma(γ) particles**
  - The third type of radioactive radiation consists of high-energy rays called **gamma(γ) rays.**
  - They have no charge and are not affected by an external electric or magnetic field.

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The Discovery of the Nucleus

Scientist: Ernest Rutherford (1911)

Experiment: Gold Foil Experiment

Setup:

- A thin gold foil was bombarded with alpha particles (positively charged helium nuclei).
- A fluorescent screen was placed around the foil to detect particle deflection.

Observations:

- Most alpha particles passed straight through the gold foil.
- Some particles were deflected at small angles.
- A few particles bounced back at large angles.

Conclusion:

- Atoms are mostly empty space.
- The nucleus is a small, dense, positively charged core at the center of the atom.
- Electrons orbit around the nucleus.

Significance:

- Rejected the "Plum Pudding Model" proposed by J.J. Thomson.
- Laid the foundation for the Rutherford Model of the Atom.
- First evidence that atoms have a central nucleus.

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The Discovery of the Neutron

- **Scientist:** James Chadwick (1932)
- **Background:**
  - The nucleus was known to contain protons, but the total atomic mass couldn't be explained by protons alone.
- **Experiment:**
  - Chadwick bombarded beryllium with alpha particles from a radioactive source.
  - The beryllium emitted an unknown radiation that could knock protons out of paraffin wax.
- **Observations:**
  - The emitted radiation had no charge and a mass similar to that of a proton.
- **Conclusion:**
  - The radiation consisted of a new subatomic particle: the neutron.
  - Neutrons are neutral particles (no charge) with a mass slightly greater than that of a proton.
- **Significance:**
  - Solved the problem of unexplained atomic mass.
  - Helped explain isotopes (atoms of the same element with different masses).
  - Crucial for the development of nuclear fission and nuclear energy.

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1.4 Make-up of the nucleus

Table 1.1: Properties of subatomic particles

Particle	Actual mass (kg)	Relative mass (amu)	Actual charge (C)	Relative charge
Proton (p)	1.672622×10^{-27}	1.007276	1.602×10^{-19}	+1
Neutron (n)	1.674927×10^{-27}	1.008665	0	0
Electron (e^-)	9.109383×10^{-31}	5.485799×10^{-4}	-1.602×0^{-19}	-1

Atomic Number and Mass Number

Atomic Number (Z):

- **Definition:** The number of protons in the nucleus of an atom.
- **Symbol:** Represented by Z.
- **Determines:**
  - The identity of an element (e.g., Hydrogen = 1, Carbon = 6).
  - The position of an element in the periodic table.

Mass Number (A):

- **Definition:** The total number of protons and neutrons in the nucleus of an atom.
- **Symbol:** Represented by A.
- **Formula:**
  - **Mass Number (A) = Protons (Z) + Neutrons (N).**
- Does Not Include: The mass of electrons (negligible).

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Atomic mass and Isotopes

- **Atomic mass**
  - **Definition:** The average mass of an atom of an element, measured in **atomic mass units (amu).**
  - Atomic Mass Unit (amu):
    - 1 amu = 1/12th the mass of a carbon-12 atom.
  - **Atomic Mass considers:**
    - Protons + Neutrons (main contributors to mass).
    - Isotopic abundance (weighted average of all isotopes).
  - Example:
    - Chlorine has two isotopes:
      - Cl-35 (75%) and Cl-37 (25%) → Atomic mass = 35.5 amu.

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- **Isotopes**
  - **Definition:** Atoms of the same element with the same atomic number but different mass numbers.
  - **Difference:** Isotopes have different numbers of neutrons.
  - **Types of Isotopes:**
    - Stable isotopes (do not decay).
    - Radioactive isotopes (decay over time, emitting radiation).
  - **Examples:**
    - Carbon-12, Carbon-13, Carbon-14 (all isotopes of Carbon).
    - **Hydrogen isotopes:**
      - **Protium** (H-1) – 1 proton, 0 neutrons.
      - **Deuterium** (H-2) – 1 proton, 1 neutron.
      - **Tritium** (H-3) – 1 proton, 2 neutrons.

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Electromagnetic radiation and Atomic spectra

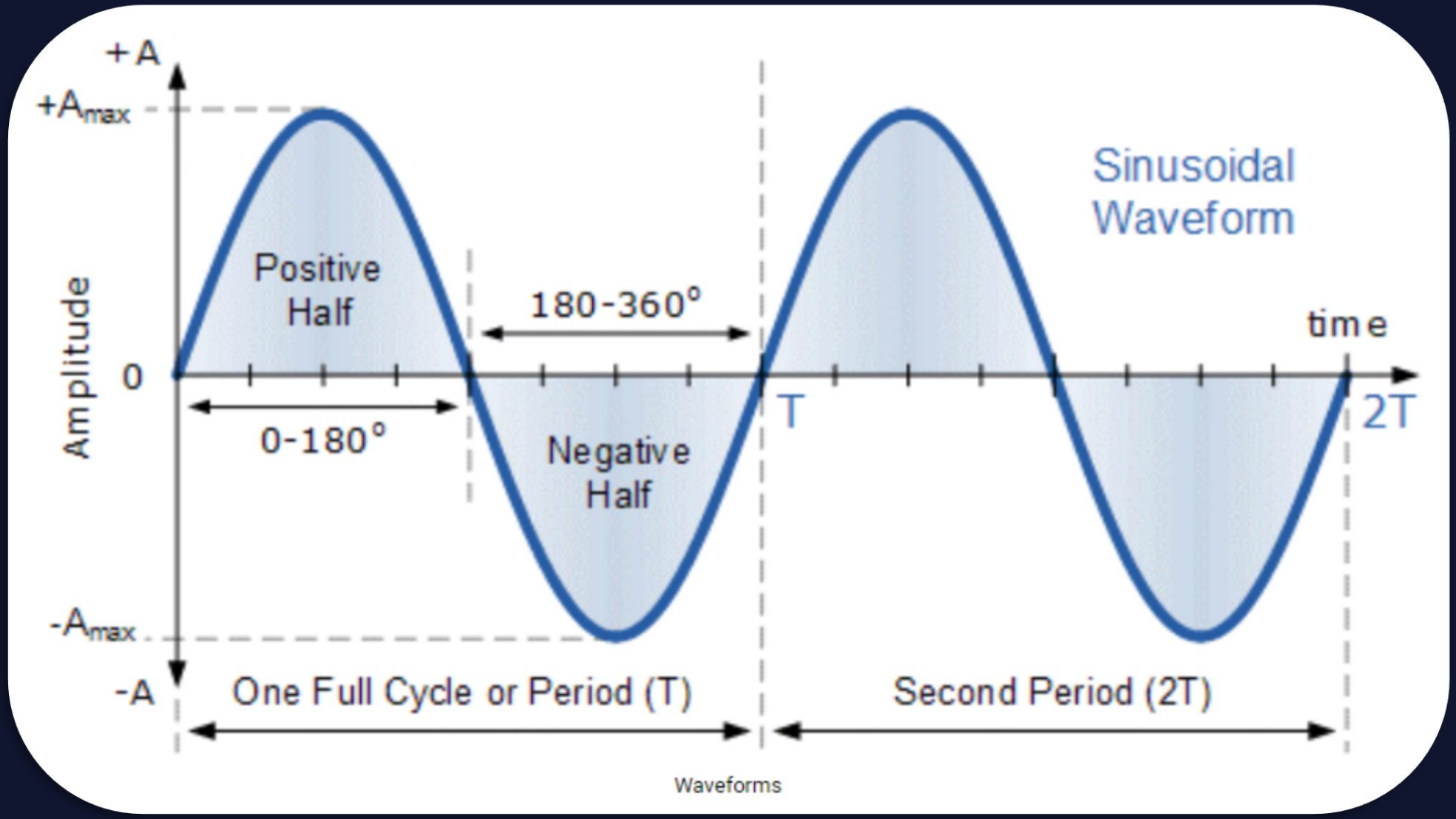
Electromagnetic radiation

- **James Clerk Maxwell's Contribution (1873)**:
  - Proposed that light consists of electromagnetic waves.
  - An electromagnetic wave has two components:
    - Electric field component.
    - Magnetic field component.
  - These components vibrate in two mutually perpendicular planes.

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- **Electromagnetic Radiation (EMR)**
  - **Definition:** EMR is a form of energy that travels through space as waves.
  - **Characterization:** Described by three properties:
    - **Wavelength (λ):** Distance between two successive crests of a wave (measured in meters).
    - **Frequency (ν):** Number of wave cycles passing a point per second (measured in Hertz, Hz).
    - **Speed of Light (c):** Constant speed in a vacuum, $c = 3.00 \times 10^8$ m/s.
  - **Relationship:** The properties are related by the formula:
    - $c = \lambda\nu$
- Increasing and Decreasing Order of Electromagnetic Spectrum
  - **Order** (increasing frequency and decreasing wavelength):
    - Radio waves < Microwaves < Infrared < Visible Light < Ultraviolet < X-rays < Gamma rays.
  - **Visible Light Range:** 400 nm (violet) to 700 nm (red)

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Quantum Theory and Photon

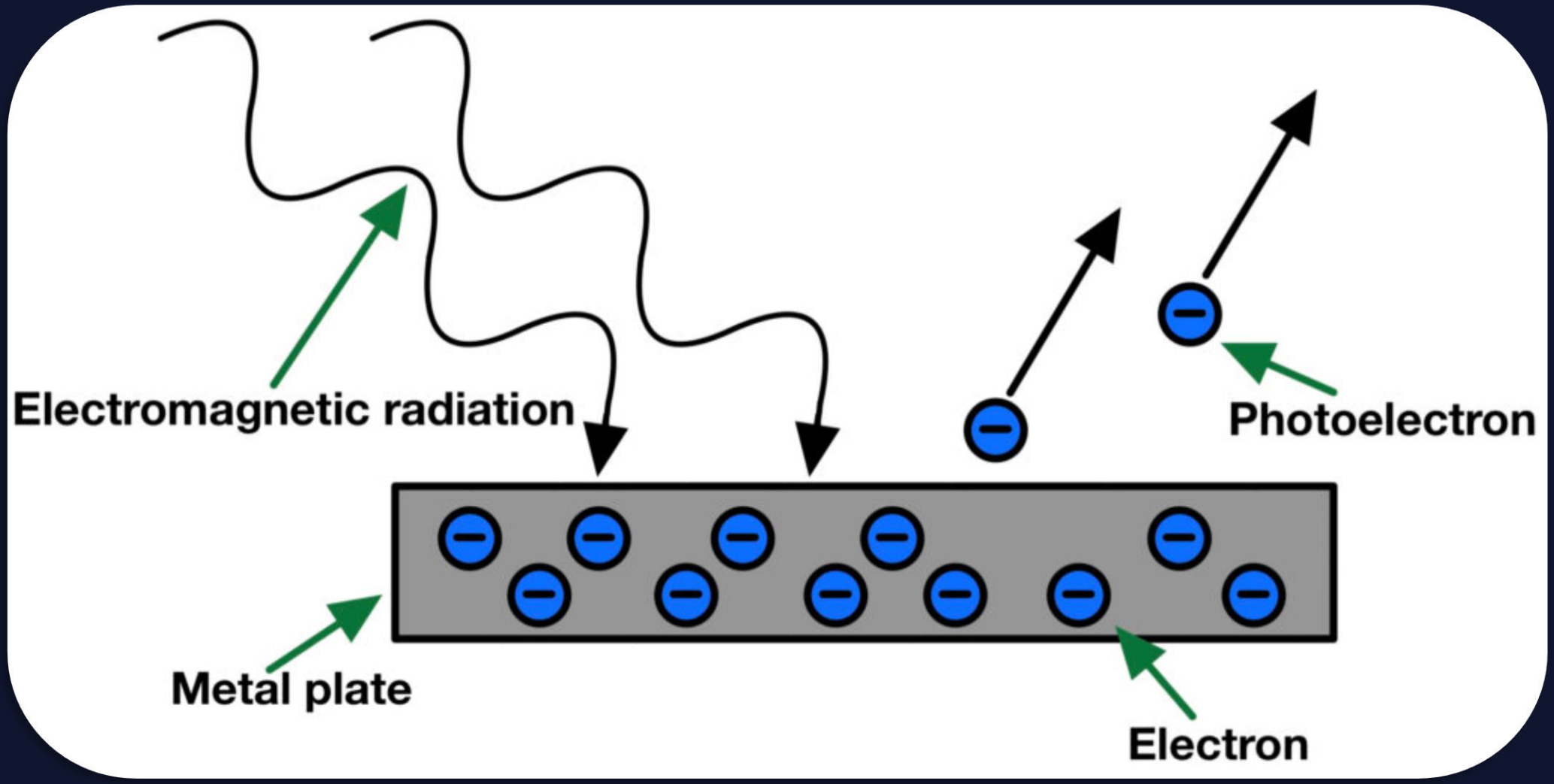
- **Proposed by Max Planck (1900):** Energy is not continuous but emitted in small discrete packets called quanta.
- **Photon:** A quantum of light energy.
- **Energy of a Photon (E):**
  - $E = h\nu$, where h is Planck's constant (6.626×10^{-34} J·s).

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Photoelectric Effect

- **Observed by:** Heinrich Hertz (1887) and explained by Albert Einstein (1905) using quantum theory.
- **Observation:** When light strikes a metal surface, electrons are ejected from the surface.
  - Photon energy must exceed the metal's work function to release an electron.
  - The kinetic energy of the ejected electron is calculated by:
    - $KE = h\nu - \Phi$, where:
      - KE = Kinetic energy of the electron.
      - $h\nu$ = Energy of the photon.
      - Φ = Work function of the metal.

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Electromagnetic radiation
Photoelectron
Metal plate
Electron

Atomic or Line Spectra

- **Definition:** Produced by the emission of photons of electromagnetic radiation (light) by excited atoms.
- **Process:**
  - An electric discharge or spark excites the gas atoms (e.g., hydrogen).
  - Energy is transferred to the electrons, raising them to higher energy states (excited states).
  - The electrons then release this energy as light when they return to lower energy states.
- **Prism Effect:** When this emitted light passes through a prism, instead of a continuous spectrum (like a rainbow), only a few specific colors (lines) are observed.
  - These lines form the element's atomic spectrum or emission spectrum.
- **Discrete Wavelengths:** Only certain wavelengths of light are emitted due to specific energy level transitions.
  - The change in energy between two levels is calculated using Planck's equation:
    - $\Delta E = h\nu = hc/\lambda$, where:

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The Bohr model of Hydrogen Atom

- **Bohr used Planck's and Einstein's ideas about quantized energy and proposed the following assumptions:**
  - **The electron in a hydrogen atom travels around the nucleus in a circular orbit.**
  - **The energy of the electron is proportional to its distance from the nucleus. The further an electron is from the nucleus, the more energy it has.**
  - **Only limited orbits with certain energies are allowed (quantized orbits).**
  - **The allowed orbits are those for which the angular momentum of the electron is an integral multiple of $h/2\pi$.**
  - **As long as an electron stays in a given orbit, it neither gains nor loses energy.**
  - **The electron moves to a higher energy orbit by absorbing light energy and emits light when falling to a lower energy orbit.**
    - **The energy (photon) of the absorbed or emitted light equals the difference between the two orbit energies.**

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- **Calculations Using Bohr's Model**
  - **Radius of Electron Orbit (r_n):**
    - $r_n = n^2 \times 0.529$ Å, where n is the orbit number.
  - **Energy of Electron (E_n):**
    - $E_n = -(13.6 \text{ eV}) / n^2$

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Hydrogen Emission Spectrum Series

- The emission spectrum of hydrogen includes a wide range of wavelengths from the infrared to the ultraviolet.
- Different spectral series are observed based on the transitions of electrons between energy level

Table 1.2: The various series in atomic hydrogen emission spectrum

Series	n_f	n_i	Spectrum region
Lyman	1	2, 3, 4, …	Ultraviolet
Balmer	2	3, 4, 5, …	Visible and ultraviolet
Paschen	3	4, 5, 6, . . .	Infrared
Brackett	4	5, 6, 7, . . .	Infrared

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Limitation of Bohr's Model

- It doesn't explain the atomic spectra of more complicated atoms and ions, even that of helium, the next simplest element.
- It doesn't explain about further splitting of spectral lines in the hydrogen spectra on application of a magnetic field.
- It considers electrons to have both known radius and orbit, which is impossible according to Heisenberg's uncertainly principle.
- Fails to explain the wave nature of electrons

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- **Dual Nature of Light**
  - Light exhibits both wave-like and particle-like behavior.
    - **Wave Nature:** Describes the interference and diffraction of light.
    - **Particle Nature:** Light consists of discrete packets of energy called **photons.**

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- **The Wave-Particle Duality of Matter and Energy**
  - **Concept Origin:** Proposed by Louis de Broglie in 1924.
  - **Main Idea:** If energy behaves like particles, then matter can exhibit wave-like properties.
  - **de Broglie's Equation:** The wavelength (λ) associated with a particle of mass m moving at velocity v is given by: $\lambda = h/mv$
    - Where:
      - λ = Wavelength of the particle (in meters).
      - h = Planck's constant (6.626×10^{-34} J·s).
      - m = Mass of the particle (in kg).
      - v = Velocity of the particle (in m/s).
  - This equation applies to all moving particles, including electrons, planets, and even larger objects. However, the wave-like properties are significant only for very small particles like electrons.

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The quantum mechanical model of the atom

- The Quantum Mechanical Model is the most advanced atomic model, developed by Erwin Schrödinger in 1926.
- This view depends on two central concepts: the wave behavior of matter and the uncertainty principle.
- Unlike earlier models, it does not describe electrons as particles orbiting the nucleus in fixed paths.
- Instead, it describes electrons as waves and their location in terms of probability clouds called orbitals.
- Orbitals are regions around the nucleus where electrons are most likely to be found.
- The model is based on Schrödinger's wave equation, which predicts the behavior and energy of electrons.
- It introduces four quantum numbers to describe the energy, shape, orientation, and spin of electrons.

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Heisenberg principle

- Heisenberg's Uncertainty Principle was proposed by Werner Heisenberg in 1927.
- It states that it is impossible to simultaneously know both the exact position and exact momentum of a particle, like an electron.
- The principle can be expressed as: $\Delta x \times \Delta p \geq h/4\pi$ where,

  Δx is the uncertainty in position,

  Δp is the uncertainty in momentum, and

  h is Planck's constant.

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- **The four quantum numbers are:**
  - The principal quantum number (n)
  - The angular momentum quantum number (ℓ)
  - The magnetic quantum number (mℓ)
  - The electron spin quantum number (ms)

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- **1. Principal Quantum Number (n)**
  - **Definition:** Represents the main energy level or shell of an electron.
  - **Values:** Positive integers (n = 1, 2, 3, ...).
  - **Significance:** Indicates the size and energy of the orbital. Higher n means a larger size and higher energy orbital.

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- **2. Angular momentum quantum number (ℓ)**
  - **Definition:** Represents the shape of the orbital.
  - **Another name:** It is also called **azimuthal quantum number** or **subsidiary quantum number** .
  - **Values:** Integers from 0 to (n − 1).
    - ℓ = 0 (s-orbital), ℓ = 1 (p-orbital), ℓ = 2 (d-orbital), ℓ = 3 (f-orbital).
  - **Significance:** Determines the shape of the electron cloud.

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3. The magnetic quantum number (mℓ)

- **Definition:** Represents the orientation of the orbital in space.
- **Another name:** Orbital - orientation quantum number .
- **Values:** Integers from −ℓ to +ℓ.
- **Significance:** Specifies the direction of the orbital in space.
- **Example:** For a p-orbital (ℓ = 1), mℓ can be −1, 0, or +1.

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- **4.The electron spin quantum number (ms)**
  - **Definition:** Represents the spin of an electron within an orbital.
  - **Values:** +½ or −½.
  - **Significance:** Describes the two possible spin states of an electron.
  - **Example:** In a paired set of electrons in an orbital, one will have ms = +½ and the other ms = −½.

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Shape of atomic orbital

Atomic orbitals represent regions around the nucleus where electrons are most likely to be found.

Each type of orbital has a distinct shape:

- **s-orbital (Spherical Shape):**
  - The s-orbitals are spherical in shape.
  - The size of the s-orbital increases with increasing energy level (e.g., 1s < 2s < 3s).
  - The electron density is uniformly distributed around the nucleus.
- **p-orbitals (Dumbbell Shape):**
  - There are three p-orbitals (px, py, pz), each oriented along one of the three axes (x, y, z).
  - They have a dumbbell or two-lobed shape with a node at the nucleus.
  - The size and energy increase with higher principal quantum numbers.

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- **d-orbitals (Complex Shape):**
  - There are five d-orbitals, and they have more complex shapes.
  - Four of them have a **cloverleaf shape**, and one (dz^2) has a **dumbbell with a ring around it**.
  - d-orbitals are important for elements in the transition metals.
- **f-orbitals (Even More Complex Shape):**
  - There are seven f-orbitals with very complex shapes.
  - These orbitals are important for heavier elements, especially in the lanthanides and actinides.

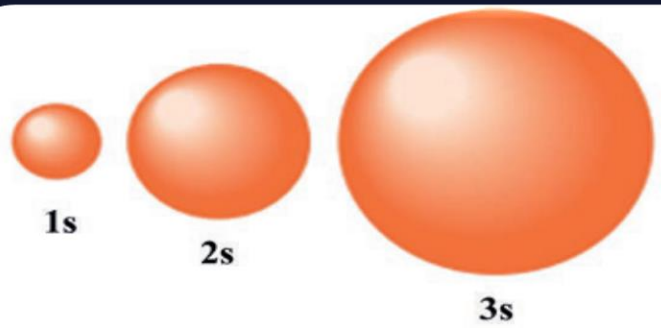


Figure 1. 15: The three s-orbitals

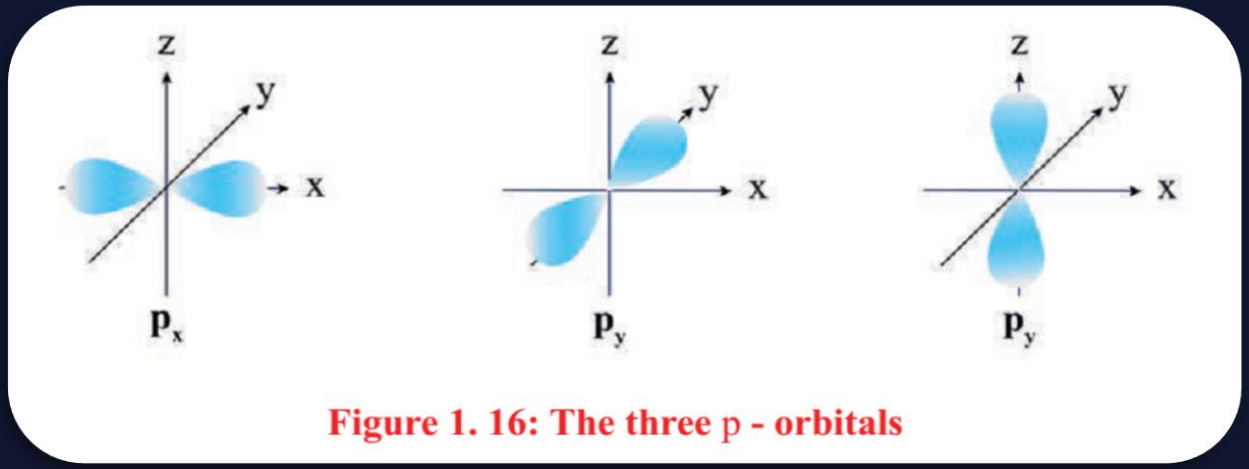


Figure 1. 16: The three p - orbitals

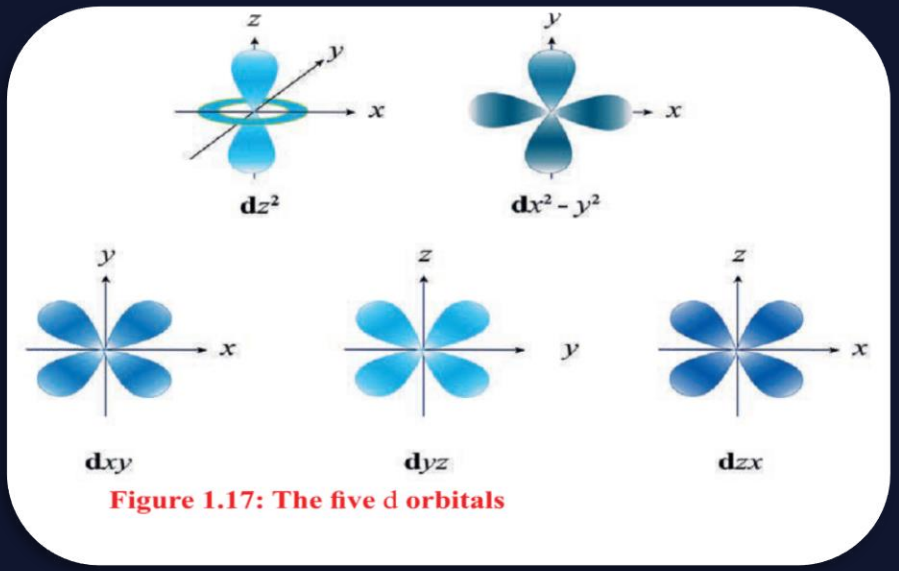
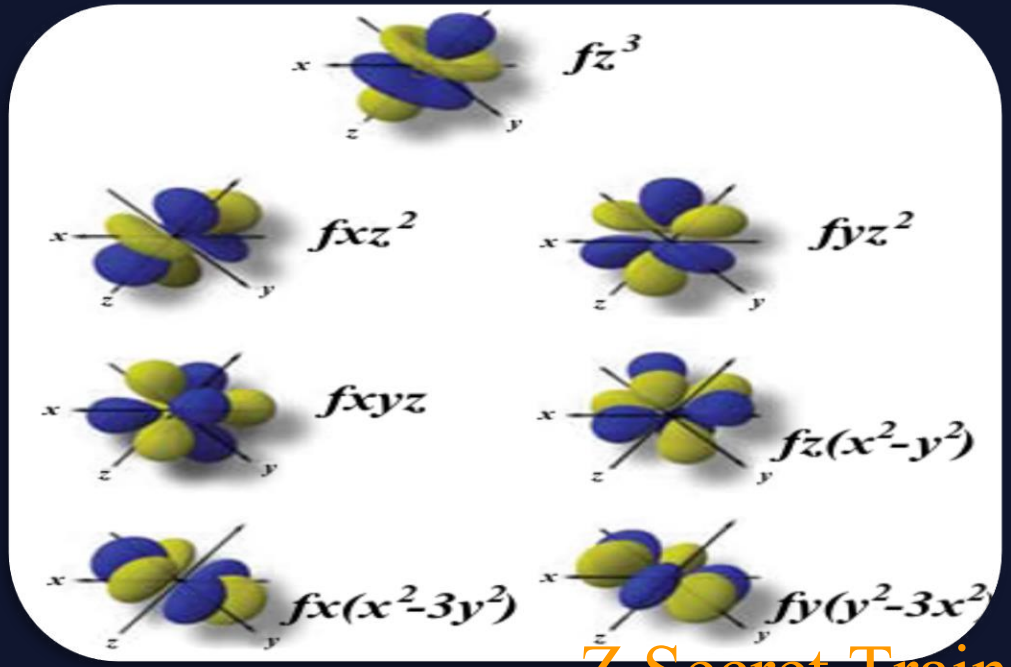


Figure 1.17: The five d orbitals



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Electron Configurations And Orbital Diagram

- **The electron configuration for any atom is governed by the following three principles:**

1. **Aufbau (building up) Principle:**
2. **Hund's Principle:**
3. **Pauli's Exclusion Principle:**

1. Aufbau (building up) Principle:

- **Definition:** The Aufbau Principle states that electrons fill atomic orbitals of the lowest available energy levels before occupying higher levels.
- **Example:** For oxygen (Z = 8), the electron configuration is $1s^2\ 2s^2\ 2p^4$.
  - The 1s orbital is filled first, then 2s, followed by 2p.

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- **2. Hund's Principle:**
  - **Definition:** Hund's Rule states that electrons will occupy degenerate orbitals (orbitals with the same energy) singly before pairing up.
  - **Significance:** Minimizes electron-electron repulsion and stabilizes the atom's electron configuration.
  - **Example:** For nitrogen (Z = 7), the 2p orbitals are filled as $2p^{-1}$, $2p^{0}$, $2p^{1}$, with each p-orbital containing one electron before pairing occurs.

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- **3. Pauli’s Exclusion Principle**
  - **Definition:** No two electrons in the same atom can have the same set of four quantum numbers.
  - **Significance:** Explains why electrons in the same orbital must have opposite spins.
  - **Example:** In a helium atom, the two electrons in the 1s orbital have opposite spins (ms = +½ and ms = –½).

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Ground state electronic configuration of the element

- **Electronic configuration** describes the arrangement of electrons in atomic orbitals within an atom.
- There are two general methods to represent electronic configurations:
  - **Subshell (sublevel) notation** – Uses numbers, letters, and superscripts to show the number of electrons in each subshell (e.g., $1s^2\ 2s^2\ 2p^6$).
  - **Orbital diagram notation** – Uses boxes or lines to represent orbitals and arrows to show electron spins (↑↓).
- The electronic configuration helps to predict an element's chemical behavior, valency, and position in the periodic table.

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Modern periodic table

- The modern periodic table arranges elements in a specific pattern based on their increasing atomic number.
- It helps to predict element properties and shows their similarities and differences.
- Periodic law states that physical and chemical properties of elements repeat periodically when elements are arranged by increasing atomic number.
- The table consists of 18 groups and 7 periods, with elements grouped by similar electron configurations.
- Elements in the same group have similar chemical properties due to the same number of valence electrons.

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Classification of the elements

- Elements are classified in the modern periodic table based on their outer-shell electron configurations, which determine their chemical behavior.
- The period number indicates the principal quantum number (n) of the outermost shell or valence shell.
- Elements are divided into metals, non-metals, and metalloids.
    - Metals are generally good conductors, malleable, and ductile.
    - Non-metals are usually poor conductors and brittle.
    - Metalloids have properties between metals and non-metals, such as boron, silicon, and germanium.

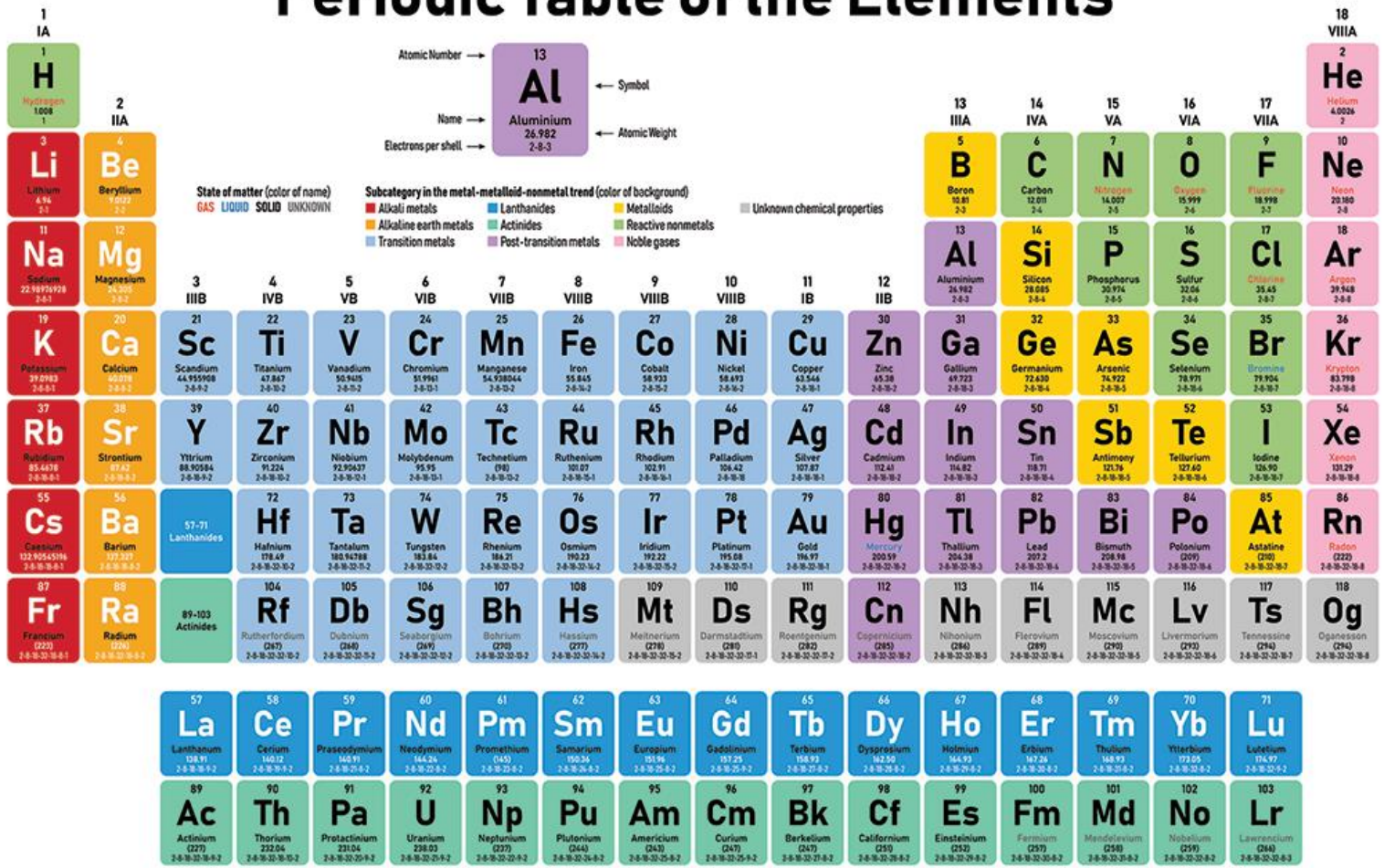
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- **Periodic Properties and Their Definitions**
  - **Atomic Radius:** The distance from the nucleus to the outermost electron.
    - **Across a Period:** Decreases due to increasing nuclear charge, which pulls electrons closer.
    - **Down a Group**: Increases due to the addition of more electron shells.
    - **Example:** The atomic radius of sodium (Na) is larger than that of chlorine (Cl).
  - **Ionization Energy:** The energy required to remove an electron from a gaseous atom.
    - **Across a Period:** Increases because of greater nuclear attraction.
    - **Down a Group:** Decreases due to increased distance and shielding effect.
    - **Example:** Helium has a much higher ionization energy than cesium.

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- **Electron Affinity:** The energy change when an electron is added to a neutral atom.
  - **Across a Period:** Becomes Increases as atoms more readily accept electrons.
  - **Down a Group:** Becomes Decreases as added electrons are farther from the nucleus.
  - **Example:** Chlorine has a high electron affinity compared to sodium.
- **Electronegativity:** The tendency of an atom to attract shared electrons in a bond.
  - **Across a Period:** Increases as atoms become more eager to attract electrons.
  - **Down a Group:** Decreases because atoms have larger radii and weaker nuclear pull.
  - **Example:** Fluorine is the most electronegative element.
- **Metallic Character:** The tendency of an element to lose electrons and form positive ions.
  - **Across a Period:** Decreases as atomic radius decreases
  - **Down a Group:** Increases as atomic radius Increases.
  - **Example:** Potassium (K) is more metallic than lithium (Li).

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Advantage of Periodic Classification of the element

- Predicting Element Properties
- Simplifying Chemical Reactions
- Understanding Trends
- Easy to Remember
- Separate Placement of Lanthanides and Actinides
- Isotopes

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THE END