Basics of the Periodic Table

1. History and Development

- **Dmitri Mendeleev (1869)**:
  - Mendeleev is credited with creating the first version of the periodic table.
  - He arranged the 63 known elements by increasing atomic mass and grouped them based on similar chemical properties.
  - His table had gaps where he predicted the existence and properties of elements yet to be discovered (e.g., scandium, gallium, and germanium).

- Mendeleev's periodic law stated: "The properties of elements are a periodic function of their atomic masses."
- **Henry Moseley (1913)**:

  - Moseley, through X-ray spectroscopy, discovered that elements are more accurately arranged by atomic number (proton number) rather than atomic mass.
  - He formulated **Moseley's Law**, which resolved anomalies in Mendeleev's table, such as the positioning of iodine and tellurium.
  - His work established the modern periodic law and led to the modern arrangement of the periodic table.

2. Periodic Law

- **Mendeleev's Periodic Law**: "The properties of the elements are periodic functions of their atomic masses."

- **Modern Periodic Law (Moseley's Periodic Law)**: "The properties of the elements are periodic functions of their atomic numbers."
    - The modern periodic table arranges elements by increasing atomic number, which determines the chemical properties of the elements.

3. Structure of the Modern Periodic Table

- **Periods**:
    - Horizontal rows in the periodic table.
    - There are 7 periods, each corresponding to the number of electron shells.
- **Groups**:
    - Vertical columns, also known as families.
    - There are 18 groups, and elements in the same group have similar chemical properties due to having the same number of valence electrons.
- **Blocks**:

- The table is divided into four blocks (s, p, d, f) based on the electron configuration of the elements.
- **s-block**: Groups 1 and 2 (alkali metals and alkaline earth metals).
- **p-block**: Groups 13 to 18 (includes metalloids, non-metals, and noble gases).
- **d-block**: Groups 3 to 12 (transition metals).
- **f-block**: Lanthanides and actinides (inner transition metals).

4. Periodic Trends

- **Atomic Radius**:
  - Decreases across a period (left to right) due to increasing nuclear charge, which pulls electrons closer to the nucleus.
  - Increases down a group as additional electron shells are added.
- **Ionization Energy**:

    - Increases across a period due to stronger attraction between the nucleus and electrons.
    - Decreases down a group because outer electrons are farther from the nucleus and more shielded by inner electrons.
- **Electronegativity**:
    - Increases across a period as the nuclear charge increases, making atoms more likely to attract electrons.
    - Decreases down a group as the distance between the nucleus and the outermost electrons increases.
- **Electron Affinity**:
    - Generally becomes more negative across a period, indicating that atoms more readily gain electrons.

    - Becomes less negative down a group due to the increased distance between the nucleus and the added electron.

5. Significance of the Periodic Table

- **Predictive Power**: The periodic table predicts the chemical and physical properties of elements and their compounds.
- **Organization**: Elements are systematically organized, making it easier to study and understand their behavior.
- **Discovery of Elements**: The periodic table has guided the discovery of new elements and the confirmation of their properties.
- **Chemical Behavior**: Helps in understanding chemical reactions and bonding patterns.