

Unit 4

PERIODIC CLASSIFICATION OF ELEMENTS

4.1 Historical Development of Periodic Classification of the Elements

4.1.1 Dobereiner's Triads

In 1829, German chemist Johann Wolfgang Döbereiner proposed the concept of triads. Döbereiner discovered that certain groups of three elements, known as triads, had similar chemical properties. Notably, the atomic mass of the middle element in each triad was approximately the average of the atomic masses of the other two elements. For example:

- **Triad 1:**
 - **Lithium (Li):** Atomic Mass = 7
 - **Sodium (Na):** Atomic Mass = 23
 - **Potassium (K):** Atomic Mass = 39
 - Average mass of Li and K = $(7 + 39) / 2 = 23$ (approximately equal to Na's atomic mass)

Although useful, this classification could not be generalized to all known elements at the time.

4.1.2 Newlands' Law of Octaves

In 1865, English chemist John Newlands observed that when elements were arranged in increasing order of their atomic masses, every eighth element had similar properties. He compared this pattern to the octave in music, where notes repeat every eight steps. This was known as the Law of Octaves. For example:

- **Octave Pattern:**
 - **Li, Be, B, C, N, O, F, Ne** (where Li and Ne show similar properties)

Limitations:

- The Law of Octaves did not work well beyond calcium (Ca).
- It did not account for noble gases, which were not yet discovered.

Activity

1. **Cultural Music and Octaves:**
 - Discuss how Ethiopian cultural music might reflect similar octave patterns as observed in Newlands' Law.

2. Classification Attempts:

- Explore the attempts by scientists to classify elements and discuss the basis for their classification.

3. Early Classification Basis:

- Explain the basis for early classification methods used by scientists.
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4.2 Mendeleev's Classification of the Elements

4.2.1 Mendeleev's Periodic Law

Dmitri Mendeleev, in the late 19th century, created a periodic table that arranged elements by increasing atomic mass. Mendeleev's periodic law stated that the properties of elements are a periodic function of their atomic masses. Key features of his table included:

- **Groups and Periods:**
 - Elements with similar properties were placed in vertical columns (groups).
 - Rows (periods) were arranged by increasing atomic mass.
- **Prediction of New Elements:**
 - Mendeleev left gaps for undiscovered elements and accurately predicted their properties.

Example of Mendeleev's Table:

- **Group I (Alkali Metals):** Li, Na, K
- **Group VII (Halogens):** F, Cl, Br

Modifications:

- After the discovery of noble gases, Mendeleev's table was expanded to include Group 0 for inert gases.
- Elements were reorganized into subgroups A and B to reflect differences in properties.

Activity

1. Construct Mendeleev's Table:

- Using the first 18 elements, arrange them in increasing order of atomic mass, draw the table, and identify the number of groups and periods.

2. Identify Patterns:

- Label the elements with their chemical symbols and observe any regular patterns.

4.3 The Modern Periodic Table

4.3.1 Modern Periodic Law

The modern periodic table, based on the modern periodic law, arranges elements by increasing atomic number rather than atomic mass. According to this law, the physical and chemical properties of elements are periodic functions of their atomic number.

Key Features:

- **Periods:**
 - There are 7 periods in the modern table, each corresponding to the number of electron shells.
- **Groups:**
 - There are 18 groups, each representing elements with similar valence electron configurations.

4.3.2 Electron Configurations

- **Periods:**
 - Elements in the same period have the same number of electron shells.
- **Groups:**
 - Elements in the same group have the same number of valence electrons.

Examples:

- **Sodium (Na):** $1s^2 2s^2 2p^6 3s^1$
- **Chlorine (Cl):** $1s^2 2s^2 2p^6 3s^2 3p^5$

Activity

1. **List Elements and Configurations:**
 - List elements from atomic number A to R and write their electronic configurations.
2. **Group Configurations:**
 - Draw vertical boxes to show elements with the same outer electron configurations. Count and record the number of sets.
3. **Period Configurations:**
 - Draw horizontal boxes to show elements with the same number of shells. Count and record the number of sets and elements per set.
4. **Patterns in Configuration:**
 - Identify and record any regular patterns observed in the arrangement.

4.4 The Major Trends in the Periodic Table

Periodic trends are the predictable patterns in properties of elements as you move across periods (rows) or down groups (columns) in the periodic table. These trends arise from the periodic nature of the elements and their atomic structures.

4.4.1 Atomic Radius

The **atomic radius** is defined as the distance from the nucleus of an atom to the outermost electron shell. There are two primary ways to determine atomic radius:

- For atoms in a metallic structure, it's half the distance between two adjacent nuclei in a crystal lattice.
- For diatomic molecules, it's half the distance between the nuclei of two bonded atoms.

Trends:

- **Within a Group:** The atomic radius increases as you move down a group. This is because additional electron shells are added, making the atom larger despite the increase in nuclear charge.
- **Across a Period:** The atomic radius decreases from left to right across a period. This happens because the increase in nuclear charge pulls the electrons closer to the nucleus, reducing the size of the atom.

Example: To compare the atomic radii of phosphorus (P), silicon (Si), and nitrogen (N), remember that N and P are in the same group, and Si and P are in the same period. Therefore, the order of increasing atomic radius is $N < P < Si$, because atomic radius increases down a group and decreases across a period.

4.4.2 Ionization Energy

Ionization energy is the energy required to remove an electron from an isolated gaseous atom. This process can be represented as: $X(g) \rightarrow X^+(g) + e^-$

Trends:

- **Within a Group:** Ionization energy decreases as you move down a group. This is because the outer electrons are further from the nucleus and are more shielded by inner electrons, making them easier to remove.
- **Across a Period:** Ionization energy increases from left to right across a period. This is due to the increase in nuclear charge, which holds the electrons more tightly and makes them harder to remove.

Exceptions: Elements in Group IIA and Group VA have higher ionization energies than their neighbors due to their stable electron configurations.

Example: For lithium (Li) and beryllium (Be), the second ionization energy for Be is higher than for Li because Be has a stable electron configuration after the first ionization, making it harder to remove an electron.

4.4.3 Electron Affinity

Electron affinity is the energy change that occurs when an electron is added to a neutral atom to form a negative ion. It can be expressed as: $X(g) + e^- \rightarrow X^-(g)$

Trends:

- **Across a Period:** Electron affinity generally becomes more negative from left to right across a period. This is because atoms are closer to filling their valence shells and thus release more energy when gaining an electron.
- **Down a Group:** Electron affinity generally becomes less negative as you move down a group. This is due to the increased distance of the added electron from the nucleus and increased electron-electron repulsion.

Example: Chlorine has a higher electron affinity than fluorine because the added electron experiences less repulsion in the larger chlorine atom compared to the smaller fluorine atom.

4.4.4 Electronegativity

Electronegativity is a measure of an atom's ability to attract and bond with electrons in a chemical bond. It is a dimensionless quantity usually measured on the Pauling scale.

Trends:

- **Across a Period:** Electronegativity increases from left to right across a period. This is due to the increased nuclear charge, which attracts electrons more strongly.
- **Down a Group:** Electronegativity decreases as you move down a group. This occurs because the increased number of electron shells reduces the nucleus's ability to attract bonding electrons.

Example: Fluorine has the highest electronegativity of any element due to its small size and high nuclear charge, while cesium has one of the lowest electronegativities because of its large atomic size and low nuclear charge.