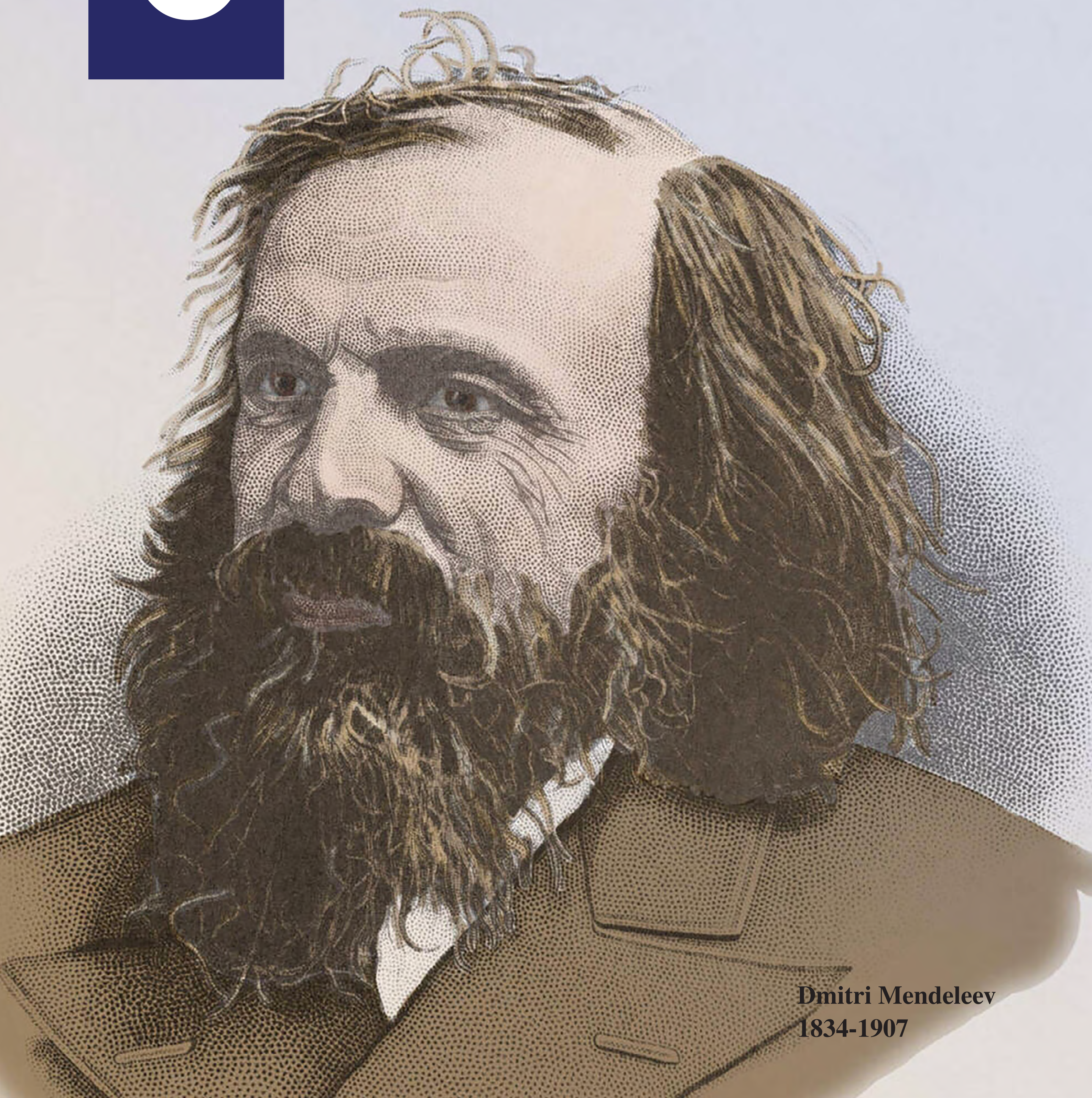
SECTION

6

PERIODIC PROPERTIES



Dmitri Mendeleev
1834-1907

SYSTEMATIC CHEMISTRY OF THE ELEMENTS

Periodicity

INTRODUCTION

In this section, you will study how periodic properties change with atomic and quantum numbers. You will describe and explain these trends, use electron configurations to determine element placement and understand how these properties vary on the periodic table.

At the end of this section, you will be able to:

- Use the electron configuration of elements to determine their position on the periodic table
- Explain how periodic properties change with atomic number and principal quantum number

Key Ideas

- **Periodicity** is the repeating pattern of chemical and physical properties of elements as you move across periods or down groups in the periodic table.
- **Periodic Table** is a tabular arrangement of chemical elements organised by increasing atomic number.
- **Row** is a horizontal line of elements in the periodic table.
- **Period** is a horizontal row in the periodic table.
- **Group** is a vertical column in the periodic table.
- **The IUPAC system** is the system of nomenclature and standardisation used by the International Union of Pure and Applied Chemistry (IUPAC).

PERIODICITY, PERIODIC TABLE AND PERIODIC LAW

Periodicity refers to the periodic or repeating trends in the chemical and physical properties of elements as they are arranged in increasing atomic numbers in the periodic table.

The periodic table is a tabular arrangement of elements organised based on their atomic structure and chemical properties. The periodic table consists of rows of elements, called periods, and columns of elements, called groups. Each element in the same group shares similar chemical properties while atoms generally decrease in size as we move across a period (due to the increasing nuclear charge attracting the electrons).

There are several types of elements in the periodic table, including metals, non-metals and metalloids (semi-metals), each with their characteristic properties and behaviour. The periodic table is an important tool and is widely used in the fields of chemistry, physics and material science.

Classification of Elements According to the Blocks

Elements are classified into four main blocks in the periodic table according to their electron configurations. These are the **s**-block, **p**-block, **d**-block and **f**-block. (*The s, p, d, and f, respectively stand for* ***s****harp,* ***p****rimary,* ***d****iffuse and* ***f****undamental*).

The classification of elements into s, p, d and f-blocks is based on the configuration of their valence electrons and their position within the periodic table. In this lesson, you will only focus on the **s**, **p** and **d**-blocks only.

1. The **s-block elements:** Elements in groups 1 and 2 are called s-block elements because their valence electrons are located in the s-subshell. These elements include metals like lithium, sodium, potassium and calcium that have low electronegativity and are highly reactive. **Helium is also a s-block element.**
2. The **p-block elements:** Elements in groups 13-18 (**except for helium**) are called p-block elements because their valence electrons are located in the p-subshell. These elements include non-metals like carbon, nitrogen and oxygen, as well as metalloids like boron and silicon. P-block elements generally exhibit a wide range of properties.
3. The **d-block elements:** Elements in groups 3-12 are called d-block elements because their valence electrons are in the d-subshell. These elements include

iron, copper and gold, and are generally harder than s-block elements. D-block elements also exhibit a wide range of properties.

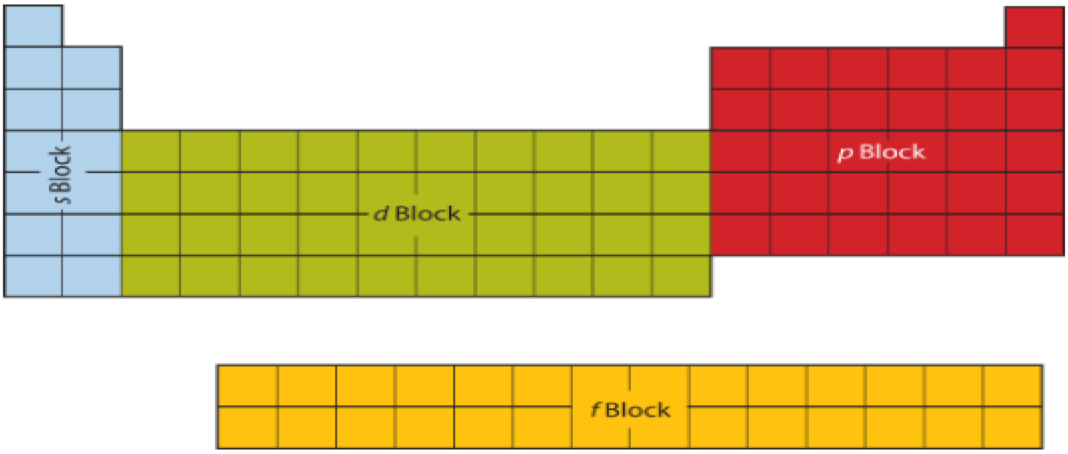


Fig 6.1: The periodic table showing the blocks. Source: (https://chem.libretexts.org/Courses/Palomar_College/PC%3A_CHEM100_-)

IUPAC and Roman Numeral Systems

The classification of elements according to groups is based on their electron configurations. The most used systems for classifying elements are the IUPAC system and the Roman numeral system.

IUPAC system

This system is used to classify elements based on the number of valence electrons they have in their outermost electron shell. The groups are numbered from 1 to 18, where group 1 contains elements with one valence electron; group 2 contains elements with two valence electrons; and so on until group 18, which contains elements with eight electrons.

Classification of elements in the IUPAC system

a. Group 1: 1 valence electron.
b. Group 2: 2 valence electrons.
c. Group 13: 3 valence electrons.
d. Group 14: 4 valence electrons
e. Group 15: 5 valence electrons.
f. Group 16: 6 valence electrons.
g. Group 17: 7 valence electrons.

h. Group 18: 8 valence electrons (**except helium**, which has 2 valence electrons).

Roman numeral system

This system is also used to classify elements based on their electron configuration. However, it assigns group numbers using Roman numerals, which are based on the number of valence electrons in the outermost shell and the subshell that contains the last electron. The numerals I to VIII are used to denote the number of outer shell electrons, and the letters A and B indicate the subshells in which the last electron is present. Here is a summary of how elements are classified using the Roman numerals:

a. Group I: 1 outer shell electron.
b. Group II: 2 outer shell electrons.
c. Group III A: 3 outer electrons in subshell A.
d. Group IV A: 4 outer electrons in subshell A.
e. Group V A: 5 outer electrons in subshell A.
f. Group VI A: 6 outer electrons in subshell A.
g. Group VII A: 7 outer electrons in subshell A.
h. Group VIII: 8 outer electrons (**except helium**, which has 2 outer electrons).

In both systems above, elements within the same group exhibit similar chemical and physical properties and have the same valence electrons in their shells.

Period of elements

The periodic table is made up of elements arranged in horizontal rows (*called periods*) from left to right, in increasing atomic numbers. The first row (*period 1*) contains only two elements, hydrogen and helium. The second row (*period 2*) contains the elements with atomic numbers 3 to 10, and the third row (*period 3*) contains elements with atomic numbers 11 through 18.

Each **period** on the periodic table corresponds to a **shell of electrons** in an atom. For example, the **first-period** elements (hydrogen and helium) have only **one shell of electrons** while the **second-period elements** have **two shells** of electrons. The number of shells increases as you go down the group. Some of the properties of elements within a period change gradually as the atomic number increases, e.g. atomic size. This is due to the increasing nuclear charge, gradual filling of electron shells and the changes in the valence electrons.

Classification of Elements According to Metals, Semi-metals and Non-metals

Elements are classified into three broad categories based on their properties. These are **metals**, **non-metals** and **semi-metals** (*also called metalloids*). The classification of elements into metals, non-metals and semi-metals is based on their physical and chemical properties and their position on the periodic table.

1. **Metals:** Metals are typically shiny, solid at room temperature (except for mercury), malleable (*ability to be hammered into different shapes*), ductile (*ability to be drawn into wires*) and good conductors of heat and electricity. They generally have high melting and boiling points and tend to lose electrons to form positively charged ions (cations). Examples of metals include sodium, aluminium, copper, iron, lead, silver and gold.
2. **Non-metals:** Non-metals are typically dull, brittle (if solid) and poor conductors of heat and electricity. They generally have low melting and boiling points and tend to gain electrons to form negatively charged ions (anions). Examples of non-metals include oxygen, nitrogen, Sulphur and carbon.
3. **Semi-metals:** Semi-metals have properties that are intermediate between metals and nonmetals. They are usually solids at room temperature, but their properties vary widely. For example, they can be shiny or dull, conductors or insulators and brittle or ductile. Examples of semi-metals include boron, silicon, Germanium, arsenic and antimony.

Physical properties of groups 1, 2 and noble gases

(*Hardness, density, melting point, boiling point and physical state*)

Group 1 (alkali metals)

a. ***Hardness**:* They are relatively soft and can be easily cut with a knife.
b. ***Density**:* They have low densities compared to most metals.
c. ***Melting point**:* They have relatively low melting points.
d. ***Boiling point:*** They have relatively low boiling points.
e. ***Physical state**:* At room temperature, all alkali metals are solid.

Group 2 (alkaline earth metals)

a. ***Hardness**:* They are harder than alkali metals but softer than most metals.
b. ***Density**:* They have higher densities than the alkali metals.

c. ***Melting point:*** They have higher melting points than alkali metals but lower than most other metals.

d. ***Boiling point:*** They have higher boiling points than the alkali metals but lower than most other metals.

e. ***Physical state:*** All alkaline earth metals are solid at room temperature.

Noble gases

a. ***Density:*** They have very low densities, particularly helium.

b. ***Melting point:*** They have very low melting points.

c. ***Boiling point:*** They have very low boiling points.

d. ***Physical state***: All noble gases are gases at room temperature except for radon, which is a radioactive solid.

Chemical properties of groups 1, 2 and noble gases

Group 1 (alkali metals)

a. ***Reactivity:*** Group 1 elements have only one electron in their outermost shell, making it easy to lose that electron in a chemical reaction. Hence, they are highly reactive and readily give up their outermost electron to form cations with a charge of +1.

b. ***Electronegativity:*** Alkali metals have low electronegativity and are highly electropositive. This makes them excellent, reducing agents in chemical reactions.

c. ***Reactivity with water:*** Alkali metals are so reactive that they can only be stored in oil or inert gases. They react vigorously with water to form hydroxides and hydrogen gas. When dropped into water, they float on the surface and release hydrogen gas with a hissing sound.

Group 2 (Alkaline earth metals)

a. ***Reactivity:*** Alkaline earth metals are reactive, but less than alkali metals. They have two electrons in their outermost shell and readily give up those two electrons to form cations with a charge of +2.

b. ***Electronegativity:*** Alkaline earth metals have low electronegativity and are highly electropositive. This makes them good, reducing agents in chemical reactions.

c. ***Reactivity with water***: Alkaline earth metals react with water, but less vigorously than alkali metals. They form hydroxides with the release of hydrogen gas.

Noble gases

Noble gases are usually chemically inactive because they have a complete outer shell of electrons. This makes them stable and non-reactive with other elements.

Activity 6.1

1. Classify the first twenty (20) elements according to the following categories:
   - **a.** The blocks (s, p, d)
   - **b.** The Groups (*IUPAC system and the Roman numeral system*).
   - **c.** The period in which the element belongs.
   - **d.** Metals, semi-metals and non-metals.
2. Explain what is meant by **s**, **p**, **d**, and **f** orbitals and give the full names of these orbitals.
3. Discuss with a friend, the physical properties (hardness, density, melting point, boiling point and physical state) of some representative elements (groups 1, 2, 7 and the noble gases).
4. Go online and watch a video on the periodic table.

   Websites:

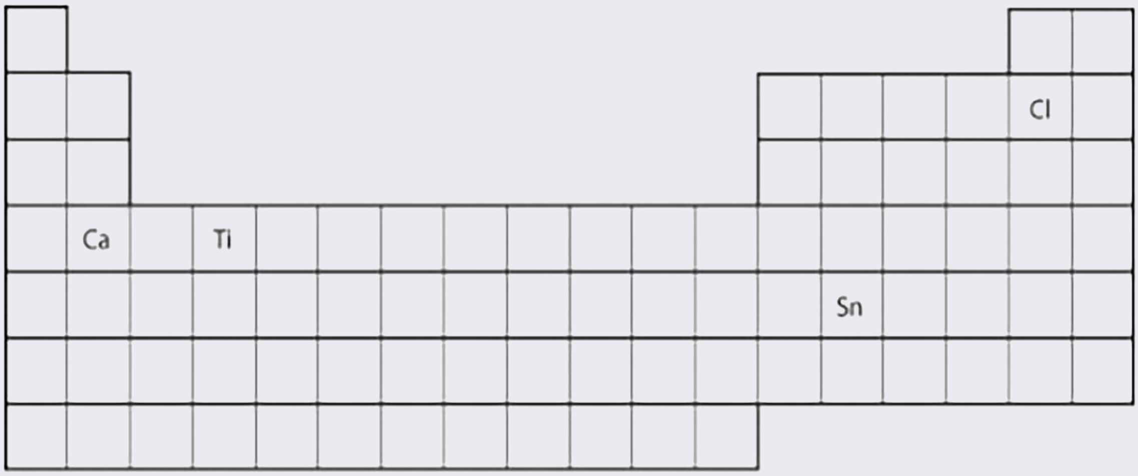
   (https://www.youtube.com/watch?v=bKKJkxqIg94)

   (https://www.youtube.com/watch?v=uPkEGAHo78o)

   (https://www.youtube.com/watch?v=wXRHz5ZEIK0)

   (https://www.youtube.com/watch?v=P6DMEgE8CK8)

5. Predict the valence shell electron configuration of the elements **Ca** and **Cl** from their positions on the periodic table below.



PERIODIC LAW

The periodic law is a statement that helps explain the repeating patterns of chemical and physical properties of elements as they are arranged in the periodic table. The periodic law states that "the properties of elements are periodic functions of their atomic numbers". In other words, elements display periodic trends in some of their properties when arranged in order of increasing atomic numbers.

This pattern arises because the properties of elements are largely determined by the number and arrangement of electrons in their atoms, and the atomic number reflects both the number of protons and electrons in an atom. The periodic table organises elements in ways that allow us to see how electrons are arranged in the outermost or valence electron shell.

Periodic Properties

A periodic property is a physical or chemical property that displays a repeating trend across the periodic table. These trends arise from the periodic variation in the electron configurations of elements.

1. **Atomic size:** Atomic size, also known as atomic radius, refers to the size of the atom. It is difficult to measure due to the uncertainty in the location of the outermost electron. It is sometimes defined as half the distance between the nuclei of two covalently bonded atoms (covalent radius).

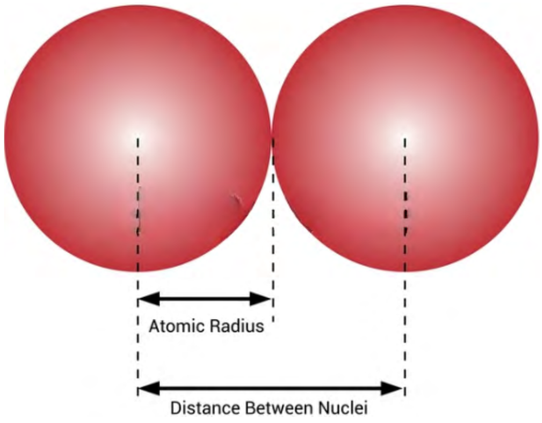


Fig 6.2: Atomic size

2. **Ionic size:** (**ionic radius:** *plural: ionic radii*). This is ***half the distance between two ions that are barely touching each other***. It is the measure of an atom's ion in a crystal lattice. It indicates the size of an ion in a crystal lattice where two atoms are bonded by an ionic bond.

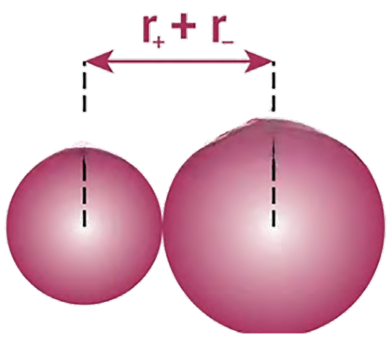


Fig 6.3: Ionic radius

3. **Ionisation energy:** The amount of energy required to remove an electron from the outermost shell of one mole of a gaseous isolated atom or molecule. There is ionization energy for each successive electron removed
4. **Electron affinity:** This is the amount of energy change that occurs when an electron is added to one mole of a gaseous neutral atom to form an anion.
5. **Electronegativity:** Electronegativity is the ability of an atom to attract electrons towards itself in a chemical bond.

Factors that affect periodic properties

1. **Nuclear charge (*atomic number*):** The number of protons in the nucleus defines the nuclear charge of an atom. The greater the nuclear charge, the greater the pull on the electrons and the smaller the atomic radii as you go

across a period. Consequently, ionisation energy and electronegativity both increase across a period on the periodic table.

2. **Distance from the nucleus (*energy level or shell*):** The distance that a valence electron is located from the nucleus is a significant factor that affects periodic properties. As the number of energy levels or shells an electron occupies increases, the electron is farther from the nucleus, decreasing the nuclear attraction to the outermost electrons.
3. **Shielding effects**: Shielding occurs when negatively charged valence electrons are shielded from attractive forces of the nucleus by inner electrons. The higher the number of inner shells of electrons, the weaker the attraction between the nucleus and the valence electron. Therefore, as the number of inner shells increases, the shielding effect increases. This effect makes it easier for outermost electrons to be removed and also shields bonding electrons from the nuclear charge, causing ionisation energy and electronegativity to decrease down a group.
4. **Electron configuration:** The electron configuration of an atom, including the number and orientation of the electrons in the valence shell, affects periodic properties. A filled valence shell or half-filled subshells contribute to stability, making these configurations highly desirable. As a result, atoms will gain or lose electrons to achieve these stable configurations.

Variations of the Periodic Property in the Periodic Table

The periodic properties show variations across the periodic table due to a variety of factors, such as ***atomic structure***, ***chemical bonding***, ***electronegativity*** and ***electron configurations***.

1. **Atomic radius:** Atomic radius refers to the distance between the nucleus and the outermost electrons. Generally, the atomic radius decreases from left to right across a period due to the increasing nuclear charge from left to right across a period, which attracts the electrons more strongly, making them more compact. This reduces the size of the atom. Conversely, atomic size increases down the group. This is because down a group, the number of energy levels (shells) increases, causing an increase in the distance between the valence electrons and the nucleus. As the number of core shells increases down the group, the shielding/screening effect increases. This decreases the nuclear attraction between the protons and the valence electrons that make the atom larger.

2. **Ionic radius:** When atoms gain or lose electrons to form ions, their sizes change. Cations (positively charged ions) are smaller than their parent atoms because they have lost electrons, which means less electron-electron repulsion and a smaller electron cloud. Anions (negatively charged ions) are larger than their parent atoms because they have gained electrons, resulting in more electron-electron repulsion and a larger electron cloud.
3. **Ionisation energy:** tends to increase from left to right across a period because there is an increase in nuclear charge which attracts the electrons more tightly, making it more difficult to be removed. Ionisation energy decreases down a group. This is because, down a group, there is an increase in atomic size and a decrease in nuclear charge which makes the valence electrons less bound to the nucleus, thereby making it easier to be removed.
4. **Electronegativity:** This typically increases as you move from left to right across a period due to an increase in the effective nuclear charge. However, there are exceptions to this trend such as boron, which has a lower electronegativity than expected due to its partially filled p-orbital. Conversely, electronegativity decreases down a group. This is because down a group, the atomic radius increases, reducing the attraction of the nucleus and bonding electrons.
5. **Electron affinity:** Electron affinity tends to increase from left to right across a period and decreases from top to bottom within a group. This is because the closer the electron is to the electrostatic attraction of the nucleus; the more energy will be released when it is added. Meanwhile, within a group, the increasing atomic size reduces the electrostatic attraction of the nucleus, making it harder to add an electron.

Discrepancies in the Periodic Properties with Respect to *Beryllium, Boron, Nitrogen And Oxygen*

1. **Size of the atoms and ions:** The size of an atom or ion affects many periodic properties. In general, atoms tend to be smaller as you go across a period due to an increase in the effective nuclear charge. However, there are some anomalies in this trend. For example, beryllium is smaller than boron, despite being to the left of it in the periodic table. This is attributed to the fact that beryllium has a fully filled 2s orbital, which makes it more compact than boron, which has a partially filled 2p orbital with greater electron-electron repulsion.

2. **Electron configuration:** The electron configuration of an element can also affect its periodic properties. For example, nitrogen has higher ionisation energy than oxygen because nitrogen has a half-filled 2p orbital, which makes it more stable than the partially filled 2p orbital of oxygen. Also, the extra electron is shielded by the half-filled 2p orbital electrons. Hence, more energy is required to remove electrons from nitrogen than is required in oxygen.
3. **Nuclear charge and shielding:** Boron has lower ionisation energy than beryllium because the 2p electron in boron is easily removed, as it experiences increased shielding from the nucleus by the filled 2s orbital. The 2s electrons in beryllium are strongly attracted by the nucleus, hence increasing the nuclear charge. Thus, more energy is needed to remove the 2s electrons of beryllium.

REVIEW QUESTIONS

1. State the periodic law.
2. Distinguish between electronegativity and electron affinity.
3. Explain why the first ionisation energy of oxygen is less than that of nitrogen.
4. Arrange the following elements in increasing order of electronegativity and explain your order: C, F, Li, Be, B, O, Na.
5. Arrange the following species in order of decreasing size K^+, Cl^- and Ca^{2+}.
6. Explain what is meant by ionisation energy and state how it varies across a period in the periodic table.
7. By using electron configuration only, find the group and period of the element with atomic number 13.
8. With examples, explain how you can predict the period and group of an element.
9. Element X has electron configuration $1s^2\ 2s^2$

a. State the block that X belongs to.

b. Which group does X belong to?

10. Consider the following elements: X and Y with atomic numbers, 17 and 11 respectively. Write the electron configuration of each element and use it to determine the:
    - block to which each belongs.
    - group to which each belongs.
    - period to which each belongs.

EXTENDED READING

Go online and log into any of these sites to read more about the elements and their daily usage in the world

- https://infinitylearn.com/surge/chemistry/first-20-elements-of-periodic-table/
- https://www.khanacademy.org/humanities/big-history-project/stars-and-elements/knowing-stars-elements/v/bhp-periodic-table-crashcourse
- https://pubchem.ncbi.nlm.nih.gov/periodic-table/

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