Periodic table

Period: The horizontal rows – there are seven periods, each of which begins with an atom having one valence electrons and ends with a complete outer shell structure.

- Magic number: 2 8 8 18 18 32

Group: The vertical columns are called Groups. There are 18 groups in periodic table. The elements in a particular Group exhibit similar properties.

Mendeleev's periodic law: 1^{st} periodic table

"The properties of elements are the periodic function of their atomic masses".

- Elements were arranged horizontally in the order of their increasing atomic masses.

Modern periodic law

The physical and chemical properties of the elements are the periodic function of their atomic number.

- Elements are arranged in order of increasing atomic numbers.

Properties related to the Periodic Table

- Elements on the left of the chart are **metals** with most reactive metals are in the lower left corner.
- **Non-metals** are found on the far right side with most active non-metals are in the upper right corner.
- The **noble or inert gases** are on the extreme right with ns^2np^6 configurations

- Group 1 metals are called **alkali metals** with ns^1 outermost configuration. They are called alkali metals since they react with water to form strong bases.

$$Na\ (s) + H_2O\ (l) = NaOH\ (aq) + H_2\ (g)$$

- Group 2 elements are called **alkaline earth metals** with ns^2 outermost configurations
- As you proceed from left to right the base-forming properties decreases and acid-forming properties increases.
- The metals in the first two Groups are **light metals**
- The metals toward the center are called **heavy metals**
- The metals along the dark line in the Periodic Table are called **metalloids**, ie. B, Si, Ge, As, Sb, Te and Po.
- Group 17 elements (or Group 7) with ns^2np^5 configuration are called **halogens group.**
- Elements of group 11 are called **coinage metals**.
- The elements in which the last electron enters the:
  s-subshell = s block elements

  p-subshell = p block elements **d**-
  subshell = d block elements (d^{1-10})
    - Ionic state (d^{1-9}) = Transitional elements

  f-subshell = f block elements = separate into two series:
    - Lanthanides & Actinides series
    - Lanthanides + Actinides series = Inner transition elements
    - Lanthanides series = Rare earth elements

Nuclear Charge

The total positive charge of all the protons in a nucleus of an atom is known as the nuclear charge.

Since the number of protons in an atom is similar to the atomic number, nuclear charge is also similar to the atomic number of the element.

The nuclear charge increases from left to right through a period and it also increase down a group.

Nuclear charge is the electrostatic force that attracts and binds the orbital electrons to the nucleus.

Effective Nuclear Charge (ENC)

The electrons in the last orbital are known as valence electrons, and they are situated farther from the nucleus.

In an atom, there is electron – electron repulsion between them. And also there is electrostatic attraction between the protons in the nuclei and the orbital electrons.

The electrons in the valence shells feel the minimum nuclear charge effect. This is because the electrons in between the nucleus and the outer shells intervene and shield the nuclear charges.

Effective nuclear charge is the nuclear charge experienced by the outer shell electrons. And this value is lower than the actual nuclear charge.

For example,

- ***Effective nuclear charge*** *= atomic number - number of nonvalence electrons*

F (9) = 1s2 2s2 2p5

ENC = 9 – 2 = 7

What is the difference between nuclear charge and effective nuclear charge?

• Nuclear charge is the total positive charge of all the protons in a nucleus of an atom. Effective nuclear charge is the nuclear charge experienced by the outer shell electrons.

• Effective nuclear charge is lower than the value of nuclear charge. (Sometimes it can be similar)

Atomic Radii *(Book: Ebbing -8.6)*

Q. What is meant by atomic radii?

It is half of the distance between the nucleus of the two atoms in a molecule, also known as co-valent radii.

- **Within a period**, the atomic radii tend to decrease with increasing atomic number. Thus the largest atom in a period is a Group IA atom, and the smallest is a noble gas.
- **Within a Group** the atomic radii tend to increase with the periodic table.

Q. How to explain this trend?

- Two factors determine the size of the outermost shells
  - (i) The larger is the *n*, the larger is the size of the orbit.
  - (ii) The effective nuclear charge (ENC) acting on an electron in the orbital. Increase of ENC reduces the size of the atomic orbital by pulling the electron inward.

Ionic Radii

It is *a measure of the size of the spherical region around the nucleus of an ion within which the electrons are most likely to be found.*

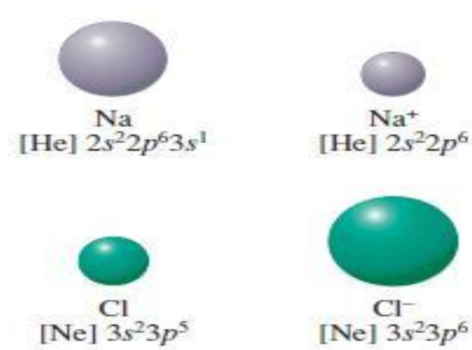

FIGURE 9.8

Comparison of atomic and ionic radii
Note that the sodium atom loses its outer shell in forming the Na^+ ion. Thus, the cation is smaller than the atom. The Cl^- ion is larger than the Cl atom, because the same nuclear charge holds a greater number of electrons less strongly.

- Ionic radii differ from atomic radii
- A Cation is smaller in size than the corresponding atom
- An anion is greater in size than the corresponding atom

Q-1. Which has the larger radius, S or S^{2-}? Explain

Q-2. Which has the smaller radius, Mg or Mg^{2+}? Explain

Explanation

A cation formed when an atom loses its valence electrons. With fewer electrons in the valence orbitals, the electron–electron repulsion is initially less, so these orbitals can shrink to increase the attraction of the electrons for the nucleus. Thus, the cation radius is smaller than the atomic radius.

Similarly, because an anion has more electrons than the atom, the electron–electron repulsion is greater, so the valence orbitals expand. Thus, the anion radius is larger than the atomic radius.

Ionization energy

The first ionization energy of an atom is *the minimum amount of energy needed to remove an electron from the outermost shell of a neutral atom in the gaseous state*.

$$Li\left(1s^{2}2s^{1}\right)\rightarrow Li^{+}(1s^{2})+e^{-}$$

The IE of Li atom is 521 $kJmol^{-1}$

Periodic variation

- **Within any Period**, IE tend to increase with atomic number
- The lowest values of IE are found for Group IA elements, while the highest values are found in the noble elements
- **Within any Group** the IE decreases with atomic number

$X(g) + energy \rightarrow X^{+}(g)\ e^{-}$ is known as **first ionization energy** (I_1).

Similarly, the second (I_2) and third ionization (I_3) are shown by the following equations:

$$X^{+}(g) + energy \rightarrow X^{2+}(g) + e^{-}\ X$$

$$^{2+}(g) + energy \rightarrow X^{3+}(g) + e^{-}$$

The ionization processes are endothermic. Energy must be supplied in order to remove the electrons from an atom in the gaseous state. The ionization energies always increase in the order---

$I_1 < I_2 < I_3 <$ ----

Explanation:

(1) When an electron is removed from an atom, the repulsion among the remaining electrons decreases

(2) Since the nuclear charge remains constant, more energy is needed to remove another electron from the positively charged ion.

Periodic variation:

(i) The IE increases along a period from left to right with the increase of atomic number. This is because of the increase in effective nuclear charge from left to right. A large effective nuclear charge means a more tightly bound electron, and hence more energy is required to remove the electron.

(ii) The IE decreases along a Group from top to bottom with the increase of atomic number.

(iii) He ($1s^2$) has the highest IE

(iv) Alkali metals (ns^1) have the lowest IE

(v) The noble gases ($1s^2$ and ns^2np^6) have highest IE

Question: Oxygen and sulfur are members of Group 6A. Which one of them has the smallest first IE?

Answer: O($2s^2 2p^4$) and S ($3s^2$ $3p^4$). The 3p electrons in sulfur atom is farther from the nucleus and experience less nuclear attraction than the 2p electrons in the oxygen atom. The first IE of O and S are 1314 and 999.5 kJ mol^{-1} respectively.

Practice questions

1. Why 2nd Ionization energy of sodium is higher than the 1st Ionization energy?
2. Why beryllium has a higher first Ionization energy than boron?
3. Why nitrogen has a higher first Ionization energy than oxygen?

Electron affinity (EA)

The amount of energy given out when an electron is added to an atom in the gaseous state.

$$X\,(g) + e^- \rightarrow X^-\,(g)$$

EA is an exothermic process.

Periodic variation of EA

(i) EA increases from left to right along a period

(ii) EA of the metals are generally lower than those of non-metals

(iii) The halogens (Group 7A) atoms have the highest EA values

(iv) Noble gases have zero EA vales

(v) Within a group the EA values vary little

Question: Why is the EA of O^- an endothermic process?

Answer: $O\,(g) + e^- \rightarrow O^-$ (g) ΔH= -141 kJ

$O^-\,(g) + e^- \rightarrow O^{2-}$ (g) ΔH= 780 kJ

After the first electron has been added to an atom, the second electron experiences more repulsion, which means more energy must be supplied in order to accommodate the second electron.

Practice questions

Why is the electron affinity of inert gases zero?

Electronegativity

It is the ability of an atom to attract toward itself the electrons in a chemical bond.

- Elements with higher electronegativity have greater tendency to attract electrons than the elements with low electronegativity.
- The electronegativity number is based on an arbitrary scale from 0 to 4. In general

Electronegativity < 2 is metal

F = 4.0	**O = 3.5**	**N = 3.5**	**Cl = 3.00**
Br = 2.7	**I = 2.5**	**C = 2.5**	**H = 2.1**

Also,

- Electronegativity decreases down a Group
- Electronegativity increase across the period.
- The lower the electronegativity number the more electropositive an element will be.
- The most electronegative element is in the upper right corner **– F**
- The most electropositive is in the lower left corner if the Periodic Table **- Fr**