

The Modern Periodic Table

The Periodic Law

- Mendeleev's periodic law stated that the properties of the elements vary periodically with their atomic masses.
- The modern Periodic Table is based on his work.
- It is essentially an arrangement of the elements in order of increasing atomic number
- The number of protons in the nucleus of an atom is given by its atomic number (Z).

The Periodic Law

- The number of protons in the nucleus of a neutral atom equals the number of electrons.
- Since similar electron configurations occur at regular intervals according to atomic number, it can be stated that it is the arrangement of the electrons in the atoms of the elements which accounts for their periodic variation in properties.

The Periodic Law

- Thus, the periodic law can be restated:
- The periodic variation in the properties of the elements occurs because the electron configurations of atoms vary periodically with atomic number

The Modern Periodic Table

- The common form of the modern Periodic Table displays the elements according to
 - increasing atomic number
 - similar outer shell electron configurations (groups)
 - order of subshell energies (blocks)
 - number of occupied shells (periods)

The Modern Periodic Table

- Remember:
 - a group is a vertical column of elements with the same outer shell electron configurations
 - a period is a horizontal row of elements with the same numbers of occupied shells (in the ground state)

Hydrogen's Position in the Periodic Table

- A hydrogen atom has only one electron and can form a positive ion (H^+) like the elements in Group 1.
- A hydrogen atom can also gain one electron to form the hydride ion (H^-) like the elements in Group VII.

Hydrogen's Position in the Periodic Table

- Hydrogen has the electron configuration $1s^1$, indicating that it is part of the s-block (Groups 1 and II), but it does not exhibit many of the usual properties of metals.
- It is therefore usually placed in a special isolated position at the top of the Periodic Table.

Glenn Seaborg (1912 - 1999)

- The Transuranic Elements
- The first transuranium element, Neptunium, (atomic number 93) was detected by American scientists in 1940 during a study of the fission products resulting from the neutron bombardment of uranium atoms.

Glenn Seaborg (1912 - 1999)

- Glenn Seaborg, also an American, identified element 94 (plutonium) in 1941 after bombarding uranium atoms with hydrogen nuclei (protons). Seaborg and his team of scientists also produced several more of these 'heavy' elements (95-103) by bombardment experiments using a particle accelerator called a cyclotron.

Glenn Seaborg (1912 - 1999)

- These elements form part of the actinide series in which the 5f orbitals are being filled.
- The transuranium elements do not occur in nature. They are classified as artificial elements because they can only be generated in a laboratory by using sophisticated equipment.
- Today, elements with atomic numbers as high as 115 have been reported. These elements are unstable (radioactive) and decay rapidly.

General Features of the Periodic Table

- Metals are found on the left hand side of each period. Properties become more metallic as atomic number increases down a group.
- Some elements, called metalloids, exhibit some metallic and non-metallic properties e.g. silicon.
- Elements which are gases at room temperature and pressure are found in the top right hand corner of the table.

General Features of the Periodic Table

- Metals in Groups I, II and III form positive Ions (cations) by losing electrons (oxidation).
- The positive charge on the ion is the same as the group number of the metal,
- e.g. aluminium:



General Features of the Periodic Table

- Non-metals in Groups V, VI and VII form negative ions (anions) by gaining electrons (reduction).
- The negative charge on the ion is the same as the difference between eight and the group number, e.g. sulfur:
 - $S + 2e^- \rightarrow S^{2-}$ **reduction**
- Elements in Group VIII (the noble gases) are usually unreactive as they have a full outer shell of electrons.

Patterns and Trends in Groups and Periods

- The Elements
- Gradual changes in the properties of elements occur down a group and across a period
- Group
 - outer shell configurations are similar
 - number of occupied shells increases
 - core charge is the same
- Period
 - Outer shell configuration change
 - Number of occupied shells stays the same
 - Core charge increases

Patterns and Trends in Groups and Periods

- The size of an atom has a profound influence on the chemical properties of an element
- The electrons in the inner shells of an atom shield the outer shell electrons from the full impact of the nuclear charge.
- Consequently, these outer shell electrons experience core charge which is less than the actual charge on the nucleus.
- Period outer shell configurations change number of occupied shells stays the same core charge increases

Core Charge

- core charge = atomic number (number of protons) - total number of electrons in inner shells only
- e.g. Group II
 - Be core charge = $4 - 2 = +2$
 - Mg core charge = $12 - 10 = +2$
 - Ca core charge = $20 - 18 = 2$
- Core charge is the same within a group

Core Charge

- In Period 3
 - sodium $11-10 = 1$
 - magnesium $12-10 = 2$
 - aluminium $13-10 = 3$
 - silicon $14-10 = 4$
- Core charge increases across a period

Atomic size

- Atomic size increases down a group because as electrons occupy most of the volume of an atom each additional occupied shell requires more room.
- F (9 electrons)
 - $1s^2 2s^2 2p^5$
 - 2 shells occupied
 - Period 2, Group VII
- Cl (17 electrons)
 - $1s^2 2s^2 2p^6 3s^2 3p^5$
 - 3 shells occupied
 - Period 3, Group VII

Atomic size

- Atomic size decreases across a period because as the charge on the nucleus increases the additional electrons go into the same shell and are attracted more strongly to the nucleus.
- in period 3
- Na ($Z = 11$) is larger than Cl ($Z = 17$)
- Na core charge +1 Group I
- Cl core charge +7 Group VII

Metallic character

- Metallic character increases down a group because as atomic size increases the outer shell electrons are less strongly bound by the nucleus and are more easily lost.
- For example, in Group 1 potassium ($Z = 19$) is more metallic than lithium ($Z = 3$).

Metallic character

- Metallic character decreases across a period because as core charge increases the outer shell electrons are more strongly bound by the nucleus and are less easily lost.
- In period 3 magnesium ($Z = 12$ electrons, core charge +2) is a metal whereas sulfur ($Z = 16$, core charge + 6) is a non-metal.

Redox Propenies

- The reducing strength of metals increases down a group.
- Positive ions are formed more easily because electrons are donated more readily.
 - For example, in Group 1 potassium ($Z = 19$), is more reactive than lithium ($Z = 3$).

Redox Propenies

- The oxidising strength of non-metals decreases down a group.
- Negative ions are formed less easily because electrons are accepted less readily.
 - For example, in Group VII bromine ($Z = 35$) is a weaker oxidant than fluorine ($Z = 9$).

Variation of Oxidising and Reducing Strength Across a Period

- Reducing strength decreases across a period
- Oxidising strength increases across a period.
- The increasing core charge reduces the ability to donate electrons.
 - For example, in period 3, sodium ($Z = 11$) Group 1, is a powerful reductant whereas chlorine ($Z = 17$), Group VII, is a strong oxidant.

Electronegativity

- measures the electron-attracting power of an atom.
- Non-metals higher than metals.
- Electronegativity tends to decrease down a group because as the number of occupied shells increases electrons are more weakly held by the nucleus.
 - For example, in Group VI, oxygen ($Z = 8$) is more electronegative than sulphur ($Z = 16$).

Electronegativity

- Electronegativity tends to increase across a period because as the core charge increases, so does the pull on the outer shell electrons which are more strongly held by the nucleus.
 - For example in period 3, sodium ($Z = 11$) is less electronegative than sulphur ($Z = 16$).

Ionisation Energy

- The first Ionisation energy *decreases down a group*
- because the core charge is constant and as atomic size increases it becomes easier to remove an outer shell electron.
- The first ionisation energy *increases across a period*
- because the core charge increases and as atomic size decreases it becomes more difficult to remove outer shell electrons.

Summary

- Atomic size, metallic character and reducing strength increase down a group and decrease across a period.
- Electronegativity, oxidising strength and first Ionisation energy decrease down a group and increase across a period.

Transition Metals

- A transition metal is an element in which the outermost s-subshell and inner d-subshell contain electrons.
- The elements of the first transition series ($Z = 21-30$) are characterised by having a partly filled 3d subshell.
- The outer 4s subshell is normally filled before the 3d.
- Atomic size decreases slightly across the first transition series, whereas the first ionisation energy and electronegativity increase slightly.

Metals

- All metals consist of a crystalline lattice containing closely packed positive ions held together by a 'sea' of mobile valence electrons.
- There are two main types of metal - the main group metals which are found in Groups I-V and the transition metals which occupy the central d-block of the Periodic Table.
- Main group and transition metals share some properties. They are both ductile, malleable, lustrous, and good conductors of heat and electricity.

Metals

- Transition metals tend to be weaker reductants (i.e. they lose electrons less readily) than main group metals.
- They generally have higher melting and boiling temperatures, are harder, more dense and of a higher tensile strength than main group metals.

Trends across the first transition series

- Nuclear charge increases across the series from scandium ($Z=21$) to zinc ($Z=30$) as electrons are added to the inner 3d-subshell.
- The outer 4s electrons are shielded by the inner 3d electrons so that atomic size decreases only slightly.
- Electronegativity and the first ionisation energy increase very little compared to the overall trend across a period.

Characteristic properties of the transition metals

- Variable oxidation states
- Both the 4s and the 3d subshell electrons can be involved in bond formation.
- Oxidation states of + 1 or + 2 occur when 4s electrons are used.
- Higher oxidation states occur when 3d electrons are used.
- Metals in Groups 1 and 11 lose their one or two outer s-subshell electrons only.

Characteristic properties of the transition metals

- Coloured compounds (except zinc)
- The photons in visible light possess enough energy to rearrange electrons within the incomplete d-subshell of transition metal ions.
- Consequently some of these photons are absorbed.
- The remaining photons are reflected or transmitted, causing the particular colour of the ion.

Formation of complex ions

In a complex Ion, a central metal cation is surrounded by several anions or polar molecules called *ligands*.

In transition metal complexes, lone pairs of electrons on the ligand form ion-dipole bonds with the central cation.

The number of ligands attached to the metal is called the *co-ordination number*.

The most common co-ordination numbers are 2,4 and 6.

Formation of complex ions

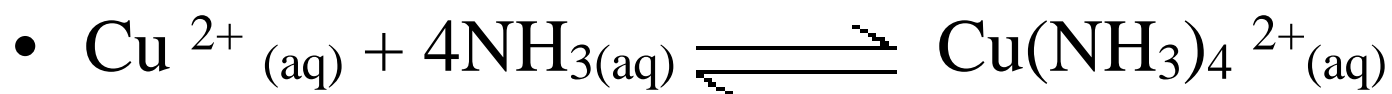
- Examples of complex ions include:
 - $\text{Cu}(\text{H}_2\text{O})_6^{2+}$ $\text{Cu}(\text{NH}_3)_4^{2+}$
 - $\text{Zn}(\text{H}_2\text{O})_4^{2+}$ $\text{Ag}(\text{NH}_3)_2^+$
 - $\text{Ni}(\text{NH}_3)_6^{2+}$ CoCl_4^{2-} ,
 - FeF_6^{3-}
- In aqueous solution, transition metal ions exist as hydrated complex ions.

Why Water and Ammonia Act as Ligands in Complexes

- Both water and ammonia are highly polar and contain lone pairs of electrons.
- This enables them to act as ligands which form ion-dipole bonds with the central metal cation:

Why Water and Ammonia Act as Ligands in Complexes

- The stability of a given complex Ion can be described by its stability constant K , which is derived from the equilibrium law expression as follows:



- $$K_{\text{st}} = \frac{[\text{Cu}(\text{NH}_3)_4^{2+}_{(\text{aq})}]}{[\text{Cu}^{2+}_{(\text{aq})}][4\text{NH}_{3(\text{aq})}]} = 1.4 \times 10^{13} \text{ M}^{-4}$$

Complex Ions

- Are important both biochemically and in industry
- Haemoglobin, the red pigment present in red blood cells, transports oxygen from the lungs throughout the body.
- The central metal cation in haemoglobin is iron (II) (Fe^{2+}) which is surrounded by four polypeptide chains.
- The complex ion, $\text{Ag}(\text{S}_2\text{O}_3)_2^{3-}$ is important in the development of photographic film.

Complex Ions

- Transition metal ions also have an important role in biological systems.
- They are required in minute amounts for the effective functioning of various enzymes (biological catalysts).
- Some important 'trace' metals are chromium, manganese, iron, cobalt, nickel and copper.
- Transition metals and their compounds can catalyse reactions because they are able to lower the activation energy of a reaction.

Magnetic properties

- Iron, cobalt and nickel are magnetic as a result of the particular arrangements of electrons in the outer shells of their atoms.

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