Chapter 10: The Periodic Table

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Designed a periodic table where he ordered elements by increasing atomic weight (now relative atomic mass). He found the chemical properties of elements recurred periodically.

Mendeleev's Periodic Law: When elements are arranged in order of increasing atomic weight (now relative atomic mass), the chemical properties of the elements recur periodically.

In this table, Mendeleev…..

- Left gaps for elements that had not been discovered yet.

- Reversed the order of some elements instead of adhering strictly to his principle of ordering them by atomic weight *(now relative atomic mass)*. E.g. he reversed the order of iodine (atomic weight 126.9) and tellurium (127.6), as it results in both elements being placed in groups that better fit their chemical properties.

Henry Moseley

In 1913, Henry Moseley discovered how to measure the atomic number of an element.

When the periodic table was **rearranged in order of atomic number**, rather than atomic mass, it **solved the problem of needing to reverse elements to fit them into specific groups – they fell naturally into place**. The modern periodic table was born.

Mendeleev's Periodic Table	Modern Periodic Table
Arranged in order of increasing atomic weight	Arranged in order of increasing atomic number
Reversed the order of some elements e.g., Te and I	Does not reverse the order of any elements
Had gaps	Has no gaps
Contained approx. 62 elements	Contains 118 elements

Modern Periodic Law: When elements are arranged in order of increasing atomic number, the physical and chemical properties of the elements recur periodically.

Positioning on the Periodic Table

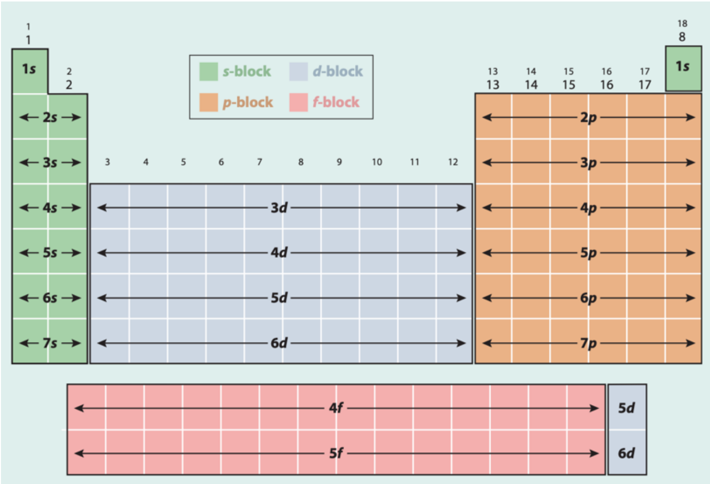
Groups: A **vertical column** on the periodic table is called a **group.** Elements in a group have **the same number of valence electrons**, and therefore have similar chemical properties.

Periods: A **horizontal row** on the periodic table is called a **period. Elements in the same period have their valence electrons in the same energy level**.

For example,,,,
In **period 1** (H & He), energy level **n=1** is occupied.
In **period 2 (**from Li to Ne), energy level **n=2** is occupied.
In **period 3** (from Na to Ar), energy level **n=3** is occupied.

A new period begins when electrons occupy a new energy level.

<u>Blocks:</u> There are **four blocks** on the periodic table called the **s-block, p-block, d-block and f-block.** These are categorised by the **atomic orbital occupied by their valence electrons**.



<u>Trends of the periodic table</u>

We will now look at some trends of the periodic table. **A trend is a general direction of data (e.g. an increase or decrease in data values).** Using the trends of the periodic table, chemists can quickly predict an element's properties.

Atomic number increases both across a period, and down a group.

Relative atomic mass values increase down a group, and usually increase across a period *(but there are some notable exceptions, e.g., Te and I).*

Metallic character decreases across a period and increases down a group. The **reactivity of metals increases down a group, while reactivity in non-metals decreases down a group.**

Electronegativity values increase across a period and decrease down a group.

<u>Atomic (covalent) radius</u>

The atomic radius (covalent radius) of an atom is defined as half the distance between the nuclei of two atoms of the same element joined by a single covalent bond.

The trend is that atomic radii values **decrease down a group and increase across a period.**

<u>First Ionisation Energy</u>

The first ionisation energy is the minimum energy required to remove the most loosely bound electron from one mole of neutral gaseous atoms in their ground state.

i.e., the energy needed for: $X_{(g)} \rightarrow X^{+}_{(g)} + e^{-}$

It is measured in kJ per mole. The lower this energy is, the more easily atoms lose electrons to become cations. The trend is that first ionisation values **decrease down a group** and **increase across a period**.

Some key terms

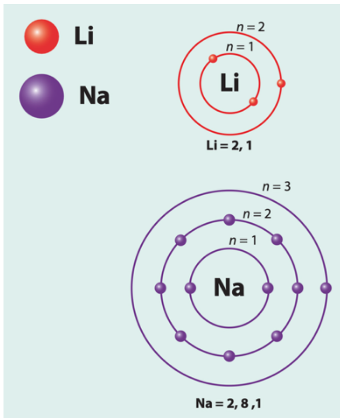
Nuclear charge is the charge of the protons in the nucleus of an atom. More protons = more charge. More charge = more influence over the negatively charged electrons in the energy levels.

Effective nuclear charge is the net positive charge experienced by an electron in a multi-electron atom, i*.e., how well the electrons can 'feel' the positive charge*. In an atom with one energy level, electrons in that energy level experience the full attractive force of the positive nucleus.

In an atom with many energy levels, the electrons in the inner energy levels shield or screen the electrons in the outer energy levels from the full effect of the nucleus. This is called **electron shielding or the screening effect.**

Trends down a group

Let's look at what happens as we move down group 1 from lithium to sodium.



We see that

- the **number of electrons increases** (from 3 to 11)
- as the number of electrons increases, **new energy levels are filled** – lithium has two occupied energy levels, while sodium has three
- there is an **increase in the screening effect**
- therefore, even though the number of protons increases, the additional screening effect means that there is **no increase in *effective* nuclear charge**

Atomic radius

Atomic radii values increase down a group for two reasons:

- **1. New energy levels occupied**
- **2. Screening effect of inner electrons**

First Ionisation Energy

First ionisation values decrease down a group (i.e. it is easier to remove an electron, turn an atom into an ion) for two reasons:

- **1. Increasing atomic radius**
- **2. Increase in screening effect**

Electronegativity

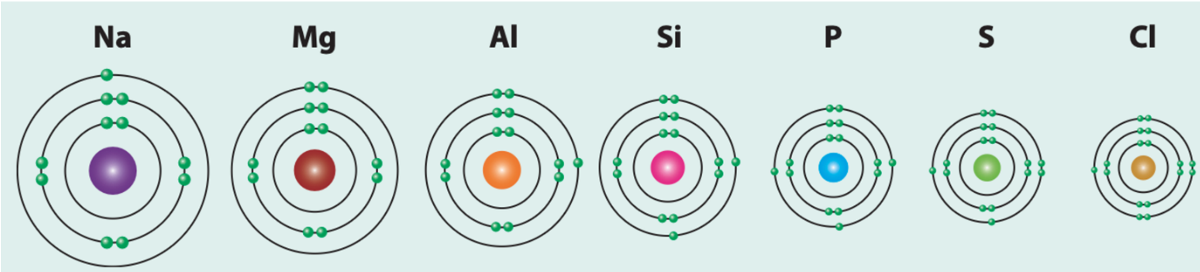
Electronegativity values **decrease down a group** (i.e. the atom is less able to attract electrons) for two reasons:

- 1. Increasing atomic radius
- 2. Increase in screening effect

Trends across a period

Let's look at what happens to an atom in period 2 as we move across from sodium to chlorine.
We see that

- **no new energy levels are filled** – sodium has three occupied energy levels, as does chlorine.
- there is **no increase in screening effect**
- the **number of protons increases** (from 11 to 17)
- **there is an increase in effective nuclear charge**



Atomic radius

Atomic radii values **decrease across a period** for two reasons:

- 1. No increase in the screening effect
- 2. Increase in effective nuclear charge

First Ionisation Energy

First ionisation energy values **increase across a period** for two reasons:

- 1. Increasing effective nuclear charge
- 2. Decreasing atomic radius

Electronegativity

Electronegativity values **increase across a period** for two reasons:

- 1. Increasing effective nuclear charge
- 2. Decreasing atomic radius

Note: These trends are for the main group elements, but you should be aware that transition metals do not follow these trends exactly.

First Ionisation Energies: Exceptions

- In the second period the values for beryllium and nitrogen are higher than expected.
- Beryllium ($1s^2\ 2s^2$) and nitrogen ($1s^2\ 2s^2\ 2p^3$) are the only two elements in the period that have either **fully filled or exactly half-filled outer sublevels**.
- Atoms with sublevels that are fully filled or exactly half-filled are **associated with extra stability**.
- **Since they are more stable (due to their fully filled or half-filled sublevels), more energy is required to remove an electron.**

(The same applies to magnesium ($1s^2\ 2s^2\ 2p^6\ 3s_2$) and phosphorus ($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$) in period three).

Successive ionisation energies

The second ionisation energy is the minimum energy required to remove a second electron from each cation in one mole of gaseous monovalent (+1) ions to give gaseous divalent (+2) ions.

i.e., the energy needed for: $X^+_{(g)} \rightarrow X^{2+}_{(g)} + e^-$

The general trend in successive ionisation energies is that they **increase as electrons are removed from an atom.**

This is due to two reasons…

1) Greater effective nuclear charge

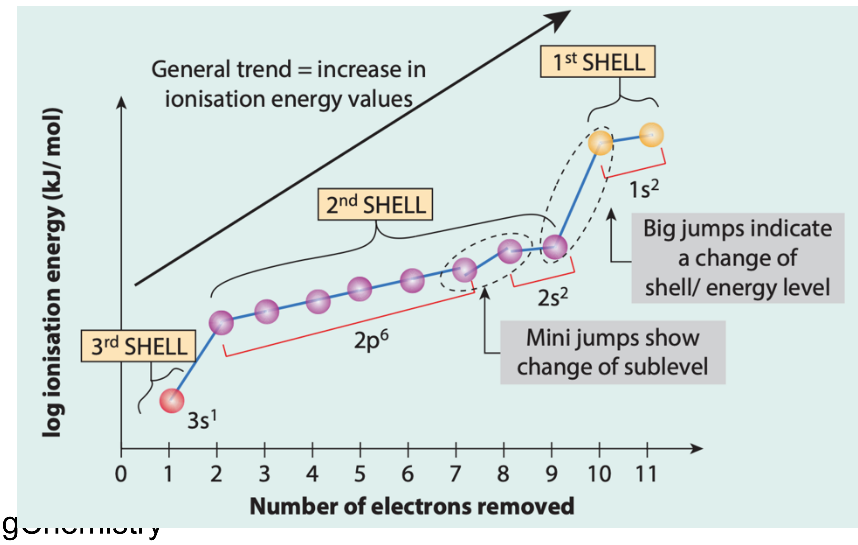
- After the first electron has been removed, all successive removals of electrons are removing electrons from a positive ion, rather than a neutral atom.
- The number of protons hasn't changed, which leads to a **greater effective nuclear charge.**

2) Decrease in radius

- This greater nuclear charge causes a decrease in radius. A **positive ion** of an element is **smaller** than the neutral atom of an element.

There are large increases in successive ionisation energies when electrons are being removed from a new energy level (which is closer to the nucleus).

We also notice a smaller (but larger than usual) increase in the energy required to remove electrons from a new sublevel. This provides evidence for the existence of sublevels (along with emission line spectra).



Properties of specific groups of the periodic table

	Group 1	Group 2	Group 7	Group 8
Name	Alkali metals	Alkaline earth metals	Halogens	Noble gases
State of matter at room temperature	Soft solids (can be cut with a knife)	Relatively soft solids	Gases *(F, Cl)* Liquids *(Br)* Solids *(I, At, Ts)*	Gases
Conductivity	Effective conductors	Conductors, but not as efficient as elements in group 1	Poor conductors	Very poor conductors
Valence electrons	1 (outer e^- in an s orbital)	2 (outer e^- in an s orbital)	7 (outer e^- in a p orbital)	8 (outer e^- in a p orbital) (ex. He: $1s^2$)
Valency	1	2	1	0
Will react to	Lose 1 electron	Lose 2 electrons	Gain 1 electron	Will not react.
Chemical Reactivity	Increases down the group (first ionisation energy values decrease)	Increases down the group (first ionisation energy values decrease)	Decreases down the group (electronegativity increases)	Unreactive as already stable (satisfy Octet Rule)