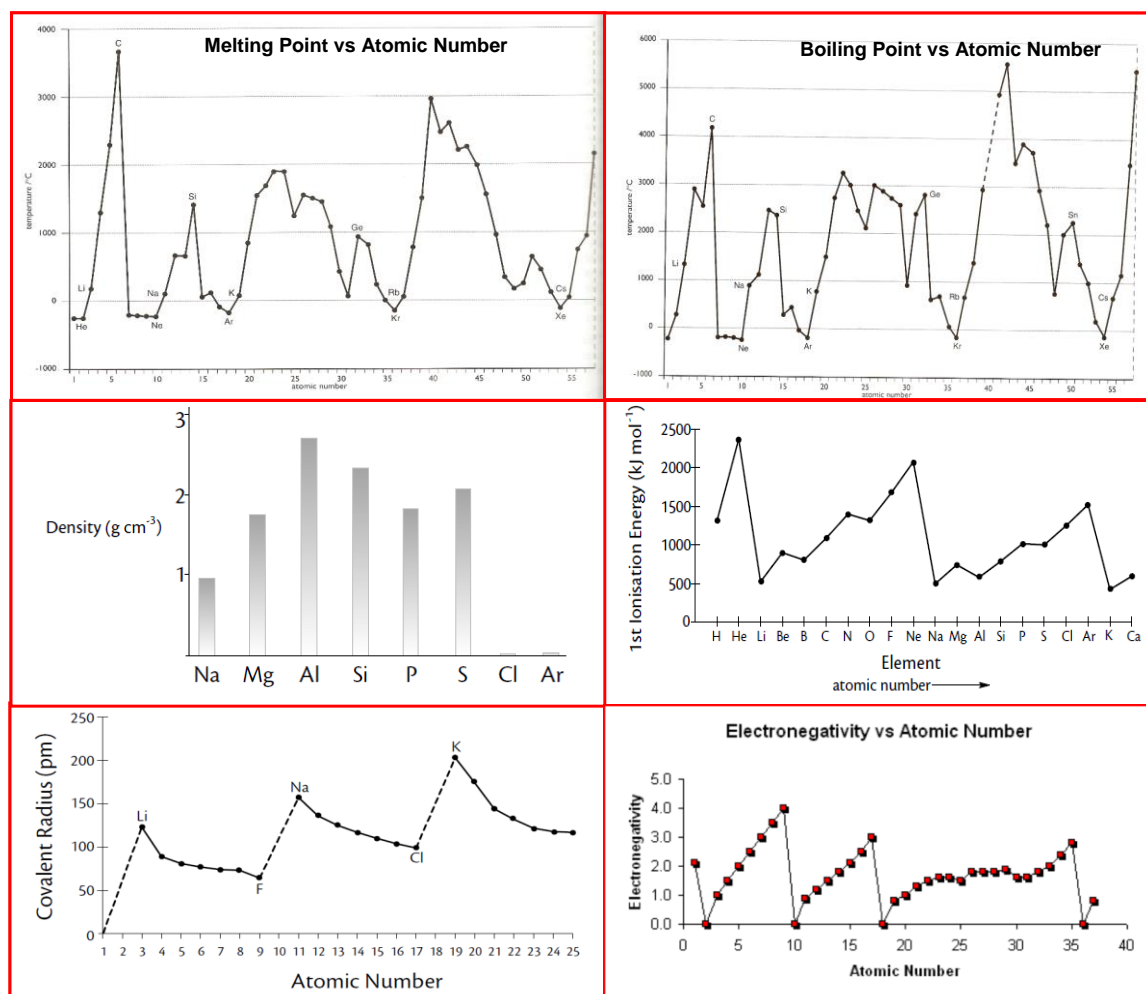# Cathkin High School
CfE Higher Chemistry

Unit 1 : Part 2
Chemical Changes & Structure
Trends in the Periodic Table

<u>Learning Outcomes – Trends in the Periodic Table</u>

Circle a face to show how much understanding you have of each statement: ☺ if you fully understand enough to do what the outcome says, ☻ if you have some understanding of the statement, and ☹ if you do not yet understand enough to do what the statement says. Once you have completed this, you will be able to tell which parts of the topic that you need to revise, by either looking at your notes again or by asking for an explanation from your teacher or classmates.

By the end of this topic I will be able to:

1. State that in the modern periodic table elements are arranged in order of increasing atomic number.   ☺ ☻ ☹

2. State that the periodic table allows chemists to make accurate predictions of physical properties and chemical behaviour for any element based on its position.   ☺ ☻ ☹

3. State that the modern Periodic Table is based on the work of Mendeleev who arranged the known elements in order of increasing atomic masses in conjunction with similar chemical properties, leaving gaps for undiscovered elements.   ☺ ☻ ☹

4. Draw graphs of properties such as first ionisation energy or covalent radius against atomic number in order to illustrate periodic trends.   ☺ ☻ ☹

☺ ☻ ☹

5. Use the data book to find the variations in the densities, melting points and boiling points of the elements across a period and down a group.

6. Explain how the melting points, boiling points and hardness / softness of elements are related to their bonding and structures.   ☺ ☺ ☹

7. Define the density of an element as its mass per unit volume, usually in g/cm^3.   ☺ ☺ ☹

8. Define the covalent radius as a measure of the size of an atom (specifically that it is half the distance between the nuclei of two bonded atoms of an element).   ☺ ☺ ☹

9. State that the atomic size decreases across a period and increases down a group.   ☺ ☺ ☹

10. Explain why there are changes in atomic size across a period and down a group.   ☺ ☺ ☹

11. Define the first ionisation energy as the energy required to remove one mole of electrons from one mole of gaseous atoms.   ☺ ☺ ☹

12. Understand that the second and subsequent ionisation energies refer to the energies required to remove further moles of electrons.   ☺ ☺ ☹

13. Explain the trends in first ionisation energy across periods and down groups in terms of atomic size, nuclear charge and the screening effect due to inner shell electrons.

☺ ☺ ☹

14. Understand that atoms of different elements have different attractions for bonding electrons.

☺ ☺ ☹

15. Define electronegativity as a measure of the attraction an atom involved in a bond has for the electrons of the bond.

☺ ☺ ☹

16. State that electronegativity values increase across a period and decrease down a group.

☺ ☺ ☹

17. Explain the trends in electronegativity across periods and down groups in terms of nuclear charge, covalent radius and the presence of "screening" inner shell electrons.

☺ ☺ ☹

Development of the Periodic Table

- The periodic table was invented by Dimitri Mendeleev (1869).
- He arranged elements in order of increasing atomic mass, and noted that their properties e.g. Melting point, boiling point and density were periodic in nature (repeating patterns existed). .
- Those elements with similar properties were placed below one another in groups and gaps were left for unknown elements.
- The **modern periodic table** is based on an elements atomic number, and this removed a number of the anomalies in the original version.
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Trends in Physical Properties of the Elements

Melting points and boiling points

- Melting points and boiling points show periodic properties. This means that they vary in a regular way or pattern depending on their position in the Periodic Table.
- Melting points and boiling points depend on the strength of forces which exist between the particles which make up a substance.
- The M.pt. & B.pt. values peak at Carbon in period 2 and at silicon in period 3.
- In general the forces of attraction (intermolecular bonding) for elements on the left of the table must be stronger, or more extensive than between the particles on the right.
- Going down group 1 the alkali metals M.pt. & B.pt. decrease so there must be a decrease in the force of attraction between the particles.

- Going down group 7 the halogens m.pt. increases so there must be a increase in the force of attraction between the particles

Density

- The density of a substance is its mass per unit volume, usually in g/cm^3.
- The density values peak at Boron (a group 3 element) in period 2 and at Aluminium (another group 3 element) in period 3.
- In general in any period of the table, density first increases from group 1 to a maximum in the centre of the period, and then decreases again towards group 0.
- Going down a group gives an overall increase in density.

Atomic Size

There is no definite edge to an atom. However, bond lengths can be worked out.
Covalent radius gives us a measure of atomic size. It is defined as half the distance between the centres (nuclei) of 2 bonded atoms. The covalent radius is measured in picometres. **pm = picometre X 10 $^{-12}$ m**
N.B. To find the bond length, add 2 covalent radii together.

- As we go across a **period**, the **nuclear charge** and the number of **outer electrons increase.**
- As we go down a **group**, the number of **electron shells** or energy levels **increases** but the number of outer electrons stays the same.
- The trends in atomic size (as measured by covalent radius) in the periodic table are:

- Across a **period** the **atomic size (covalent radius) decreases** as the nuclear charge increases and attracts the outer electrons closer to the nucleus.
- **Down a group** the **atomic size (covalent radius) increases** as an extra electron shell is added.

First Ionisation Energy

This is defined as *"the amount of energy required to remove one mole of electrons from one mole of atoms in the gaseous state"*

$M_{(g)} \rightarrow M^+_{(g)} + e$ 1st ionisation
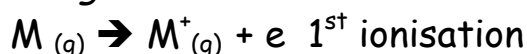
The outermost electron will be the most weakly held and is removed first.

The ionisation energy is an enthalpy change and therefore is measured per mole.

Its units are $kJmol^{-1}$ (kilojoules per mole).

Second Ionisation Energy

This is defined as *"the amount of energy required to remove one mole of electrons from one mole of gaseous 1+ ions"*

$M_{(g)}^+ \rightarrow M_{(g)}^{2+} + e$ 2nd ionisation
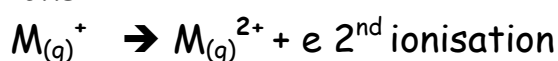
Trends in Ionisation energy

- *Across a period ionisation energies increase.*
  - ❖ This is because the nuclear charge increases (greater positive charge on the nucleus) and holds the outer electrons more strongly. More energy needs to be supplied to remove the electron.
  - ❖ Within each period the noble gas has the highest value for the 1st ionisation energy explaining the stability of full electron shells.

- ***<u>Down a group ionisation energies decrease.</u>***
  - ❖ This is because the outer electrons are further away from the nucleus. The screening effect of the inner electron shells reduces the nuclear attraction for the outer electrons, despite the increased (positive) nuclear charge.

<u>Why is the second ionisation energy of an element always greater than the first ionisation energy?</u>

In the second ionisation energy negative electrons are being removed from positive ions rather than neutral atoms.

In the positive ion there is a greater attraction for the electrons so more energy is needed to remove the second mole of electrons.

<u>Why is the second ionisation energy of K much greater than the second ionisation energy of Mg?</u>

The second ionisation of K involves removing an electron from a stable electron arrangement.

This requires a lot of energy.

The second ionisation of Mg involves removing an electron to form a stable electron arrangement.

This requires less energy.

<u>Trends in Ionisation energy</u>

- Successive ionisation energies increase as the atom becomes more positive.
- There is a large jump in ionisation energy when the electron to be removed comes from a new shell, closer to the nucleus.

  e.g. between the 2nd and 3rd ionisation energy for magnesium.

- The total energy to remove more than 1 mole of electrons is equal to the sum of each mole added together (as above).

Electronegativity

Electronegativity is a measure of an atom's attraction for the shared pair of electrons in a bond

Which atom would have a greater attraction for the electrons in this bond and why?

Linus Pauling

Linus Pauling, an American chemist (and winner of two Nobel prizes!) came up with the concept of electronegativity in 1932 to help explain the nature of chemical bonds.

Since fluorine is the most electronegative element (has the greatest attraction for the bonding electrons) he assigned it a value and compared all other elements to fluorine.

Today we still measure electronegativities of elements using the Pauling scale.

Values for electronegativity can be found on page 10 of the data book

Electronegativities

In the element chlorine both atoms have the same electronegativity so the electrons are shared equally.

In the compound hydrogen iodide the bonded atoms have different electronegativities. The iodine atom has a bigger attraction for the shared electrons than the hydrogen atom. As the electrons are attracted closer to the iodine it becomes slightly negative (δ-) and the hydrogen atom becomes slightly positive (δ+).

The trend in electronegativity is:
- **Across** a period, **electronegativity increases**.
- This is because the nuclear charge increases, attracting the electrons more strongly to the nucleus. As a result, the electronegativity increases.
  - **Down** a group, **electronegativity decreases**.
- Going down the group, the nuclear charge increases but the number of electron shells also increases. As a result of 'shielding' and the increased distance the outer shell is from the nucleus, electronegativity decreases.

Trends in the Periodic Table - Glossary

Word	Meaning
Covalent radius	Half the distance between the nuclei of two bonded atoms of an element.
Density	The density of a substance is its mass per unit volume, usually in g/cm3.
Electronegativity	A measure of the attraction that an atom involved in a bond has for the electrons of the bond.
First ionisation energy	The amount of energy required to remove one mole of electrons from one mole of atoms in the gaseous state.
Periodicity	The occurrence of patterns in the Periodic Table.
Screening effect	When inner electrons shield an outer electron from the attractive effect of the nucleus and less energy is needed to remove the outer electron as a result.