WJEC Wales Chemistry GCSE

1.2: Atomic Structure and the Periodic Table

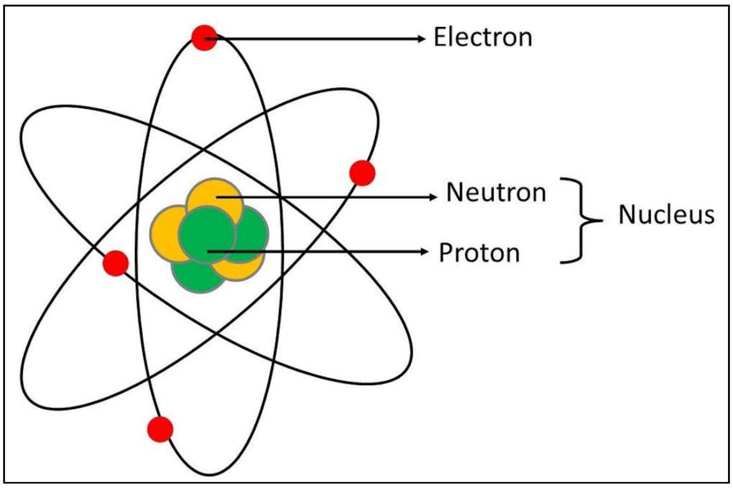
Detailed notes

Structure of the Atom

The particles that make up an atom are called subatomic particles.

There are 3 types:

- Protons
- Neutrons
- Electrons



The protons and neutrons make up the nucleus which contains most of the mass of an atom and is found in its centre. Protons are positively charged and neutrons have no net charge, giving the nucleus an overall positive charge. Orbiting the nucleus is the smallest subatomic particle, the electrons, which are negatively charged.

Subatomic Particle	Charge	Relative mass
Proton	+1	1
Neutron	0	1
Electron	-1	1/2000

Atoms have no overall charge because the number of protons and electrons are equal. Ions are formed when the number of protons and electrons aren't equal; this always results from the loss or gain of electrons, not protons.

- Atomic number - The number of protons an atom has dictates what element it is.
- Mass number - The relative mass of the atom, given by the total number of protons and neutrons.
- Atoms of the same element (therefore the same number of protons) can have different numbers of neutrons; these atoms are called isotopes of that element and have different masses.

<u>HIGHER TIER - calculating the relative atomic mass of elements with more than one isotope</u>
The relative atomic masses given in the periodic table often have decimals. This is because the average mass of that particular element is given, taking into account the abundance of different isotopes. The relative atomic mass can be calculated by using the following equation:

$$\text{Relative atomic mass} = \frac{(\%\text{ abundance of isotope 1 x mass of isotope 1}) + (\%\text{ abundance of isotope 2 x mass of isotope 2}) + \text{etc. for all isotopes}}{100}$$

Example question - carbon has 2 isotopes: carbon-14 with abundance 20% and carbon-12 with abundance 80%. Calculate the relative atomic mass of carbon.

To calculate it: ((isotope 1 mass x abundance) + (isotope 2 mass x abundance)) ÷ 100

For this question: ((14 x 20) + (12 x 80)) ÷ 100 = 1240 ÷ 100
= 12.4

The periodic table

Overview

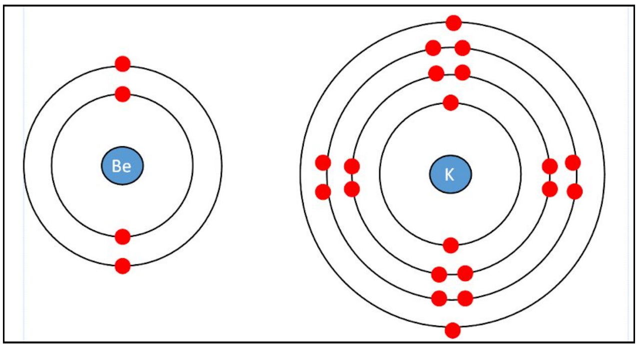
- Elements are arranged in order of atomic (proton) number which places elements with similar properties in columns, known as groups.
- Elements in the same periodic group have the same amount of electrons in their outer shell, which gives them similar chemical properties.

Organisation of elements

- Metals are found on the left side and centre of the periodic table.
- Non-metals are on the right side.
- Going across the period (row), elements with intermediate properties are found between non-metals and metals.

Electron configurations

- Electrons are organised in shells around the nucleus.
- Electrons occupy the lowest available energy levels, the shells closest to the central nucleus.
- The lowest energy shells must be filled before adding electrons to other shells.
- The first shell can hold 2 electrons, the next 2 shells can each hold 8 electrons and for the 4th shell you will only ever have to fill up to a final 2 electrons.
    - 2,8,8,2 is the largest electron configuration you can be tested on, this is a total of 20 electrons so is for the element calcium.
- For an atom, the number of protons is equal to the number of electrons so the atomic number tells you how many electrons to put into shells.
- Examples:
    - Be (4 electrons): 2,2
    - P (15 electrons): 2,8,5
    - K (19 electrons): 2,8,8,1
- These electron configurations can be represented diagrammatically:



- The electron configuration of an atom is related to its location in the periodic table:
  - The group number an atom is in (e.g. group 1) is equal to the number of electrons in its outer shell.
  - The period is related to the total number of electron shells occupied.

Trends in the periodic table

Elements in the same group have similar chemical and physical properties and this is best illustrated by groups 1 and 7. Electron configuration determines chemical properties so groups have similar properties as they have the same number of electrons in their outer electron shell.

Ion formation of group 1 and group 7

- Many reactions of these 2 groups involve the loss or gain of electrons, forming ions.
- Group 1 and 7 elements lose and gain electrons to form a full outer shell of electrons, which is more stable than an unfilled shell:
  - Group 1 elements lose one electron to form a +1 ion.
  - Group 7 elements gain one electron to form a -1 ion.

HIGHER TIER ONLY:

- **Reactivity of group 7 decreases down the group because:**
  - **Halogens react by gaining an electron (to increase their number of outer shell electrons from 7 to 8).**
  - **The number of shells of electrons increases down the group, so down the group the element attracts electrons from other atoms less, so can't react as easily.**
- **Reactivity of group 1 increases down the group because:**
  - **Alkali metals react by losing an electron.**
  - **The number of shells of electrons increases down the group. Electrons further from the nucleus are held to it less strongly so are lost more easily, making larger group 1 atoms more reactive.**

Group 1 - alkali metals

- They have characteristic properties due to the single electron in their outer shell.
- Metals in group one react vigorously with water to create an alkaline solution and hydrogen.
- They all react with oxygen to create an oxide.
- They all react with chlorine to form a white precipitate.
- The reactivity of the elements increases going down the group:

	Reaction with oxygen	Reaction with water	Reaction with chlorine
Lithium	Burns with a strongly red-tinged flame and produces a white solid	Fizzes steadily, gradually disappears	White powder is produced and settles on the sides of the container
Sodium	Strong orange flame and produces white solid	Fizzes rapidly, melts into a ball and disappears quickly	Burns with a bright yellow flame, clouds of white powder are produced and settles on the sides of the container
Potassium	Large pieces produce lilac flame, smaller ones make solid immediately	Ignites with sparks and a lilac flame, disappears very quickly	Reaction is even more vigorous than with sodium

Group 7 – The halogens

- Similar reactions due to their seven electrons in their outer shell.
- Non-metals-exist as diatomic molecules made of pairs of atoms.
  E.g. Cl_2
- They react with metals to form ionic compounds in which the halide ion carries a -1 charge.
- They react with non-metals to form covalent compounds, where there is a shared pair of electrons.
- As you go down the group, relative molecular mass, melting point and boiling point all increase.

HIGHER TIER ONLY:

Decrease in reactivity means that a more reactive halogen (one from higher up group 7) can displace a less reactive one in an aqueous solution of its salt.

- **E.g. Chlorine will displace bromine if we bubble the gas through a solution of potassium bromide:**

  Chlorine + Potassium Bromide → Potassium Chloride + Bromine
- **Chlorine will displace bromine and iodine**
- **Bromine will displace iodine but not chlorine**
- **Iodine can replace neither chlorine or iodine**

<u>**Uses of chlorine**</u>

- Chlorine is a disinfectant and kills bacteria so is used to sterilise drinking water and clean swimming pools.
- Reacts with sodium hydroxide and water to form bleach.

- Used in the manufacture of chemicals including insecticides, PVC (as polymers) and chlorofluorocarbons.

Uses of iodine

- Iodine is an antiseptic so can be used to prevent infection in hospital procedures.

The reactions of halogens with alkali metals and with iron

The halogens all react quickly with alkali metals to form a crystalline halide salt. For example, the reactions of potassium with each of the halogens are:

- $2K + F_2 \rightarrow 2KF$ (potassium fluoride)
- $2K + Cl_2 \rightarrow 2KCl$ (potassium chloride)
- $2K + Br_2 \rightarrow 2KBr$ (potassium bromide)
- $2K + I_2 \rightarrow 2KI$ (potassium iodide)

Halogen	Reaction with iron wool observations
Fluorine	Cold iron wool reacts almost instantly to form white iron(III) fluoride.
Chlorine	Reacts vigorously to form an orange-brown precipitate of iron chloride.
Bromine	Reacts quickly to form a red-brown precipitate of iron bromide. The reaction has to be warmed.
Iodine	Reacts slowly in iodine vapour to form a grey iron iodide precipitate. The reaction has to be heated strongly.

Group 0

Group 0 elements are chemically inert compared to other elements

- They have 8 electrons in their outer shell (except helium, which has 2- but this shell is still full).
- Their full outer shell makes them unreactive because they are very stable.
- They are monatomic.

Uses of group 0 elements

- Helium
  - Helium has a very low density so it is used in balloons and airships since it is much less dense than air, so balloons filled with it float upwards.
- Argon
  - Is very inert and non-flammable so is used inside light bulbs and stops the filament burning away.
  - Used as a shield gas during welding due to its inertness.
- Neon
  - Used in advertising signs; it glows when electricity is passed through it and different coloured glows can be created by coating the glass tubing with other chemicals.

Chemical tests

Flame tests to identify metal ions

Method:

1. Clean a metal loop by dipping it in hydrochloric acid then holding it in a Bunsen burner blue flame.
2. Dip the loop into the test sample.
3. Hold the loop in a blue Bunsen burner flame and observe the colour.
4. Repeat 1-3 for the next sample.

Metal ion	Colour
Li^+	Red
Na^+	Orange-yellow
K^+	Lilac
Ca^{2+}	Orange-red
Ba^{2+}	Green

Test with silver nitrate solution to identify halide ions

Method:

1. Add a few drops of dilute nitric acid to the sample; the nitric acid will react with any carbonate ions present. Carbonate ions produce a white precipitate when reacted with silver nitrate solution, so could mask the result of any halide precipitates and would give the same coloured precipitate as chloride ions.
2. Then add a few drops of dilute silver nitrate solution.
3. Observe the colour of any precipitates formed.

	Chloride, Cl^-	Bromide, Br^-	Iodide, I^-
Precipitate colour	White	Cream	Yellow
Ionic equation	$Ag^+ (aq) + Cl^- (aq) \rightarrow AgCl (s)$	$Ag^+ (aq) + Br^-(aq) \rightarrow AgBr (s)$	$Ag^+ (aq) + I^-(aq) \rightarrow AgI (s)$

Test for hydrogen gas

Method:

1. If you suspect a gas being released from a reaction is hydrogen then collect some of the gas in an upturned test tube over the reaction mixture.
2. Place a lighted splint into the test tube.
3. A squeaky pop sound means hydrogen gas is present.