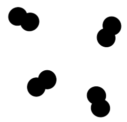
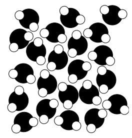
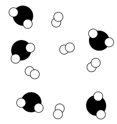
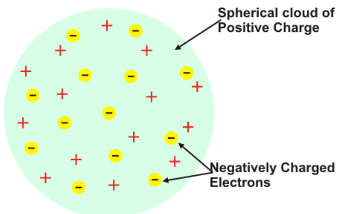
# AQA Trilogy-Chemistry key terms – Atomic Structure and the periodic table

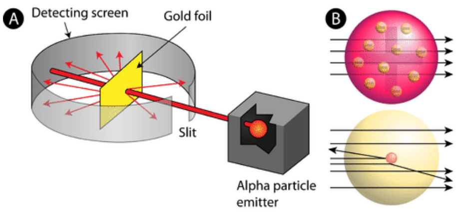
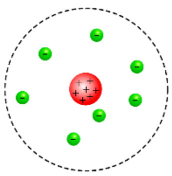
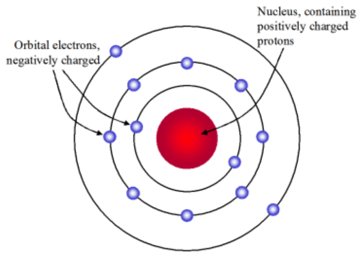
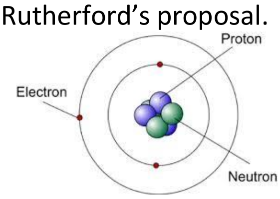
## Key ideas from KS3

| **Elements**: | **Compounds**: | **Mixtures**: |
|---|---|---|
| • are made of particles called atoms<br>• contain only one type of atom – meaning they cannot be broken down into simpler substances.<br>• have their own symbols and are listed in the periodic table<br>• are either metals or non-metals<br> | • are substances made from two or more *different* elements chemically joined (i.e. bonded) together.<br>• have different properties from the elements from which they are made<br>• are difficult to break back down into their elements<br> | • Are substances that are *not* chemically joined together<br>• Can be easily separated by a range of techniques, including:<br>✓ Filtration<br>✓ Evaporation<br>✓ Crystallisation<br>✓ Distillation<br>✓ Fractional distillation<br>✓ Separating funnel<br>✓ chromatography<br> |

## Atomic Structure

Ideas about atoms have changed over time, with an increase in knowledge and technology. New evidence has been gathered from the experiments of scientists who used their model of the atom to explain their observations and calculations. The most important advancements are listed below.

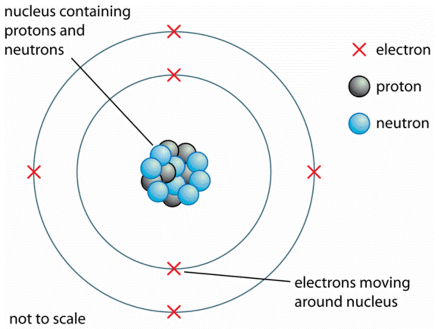
| John **Dalton**<br>(early 1800s) | J.J. **Thompson**<br>(late 1800s) |
|---|---|
| • Used experiments to suggest **substances were made up of tiny spheres called atoms**, which were the fundamental building blocks of nature.<br>• He also suggested **that chemical elements each had their own atoms**, which differed from others by mass.<br>• **Discovered the electron** by applying high voltages to gasses at low pressure. | • Suggested the **'plum pudding' model of atoms** – tiny negatively charged electrons embedded in a cloud of positive charge.<br>• As atoms are neutral, the number electrons and positive charge must be equal.<br> |

| Ernest **Rutherford**<br>(1909) | Niels **Bohr**<br>(1914) | James **Chadwick**<br>(1932) |
|---|---|---|
| Based his suggestions on the Gold Foil / Alpha particle Experiment conducted by Geiger and Marsden…<br><br>The positively charged alpha particles were shot at gold foil and expected to go straight through, but actually scattered (i.e. deflected).<br><br>Suggested that the positive charge (**protons**) are found concentrated in a central part of the atom, its **nucleus**. This is the **nuclear model of atoms**.<br> | Suggested electrons orbit the nucleus at set distances (i.e. **energy levels**)<br><br>Bohr's theoretical calculations agreed with experimental observations. | Discovered the **neutron**. This supported Rutherford's proposal.<br> |

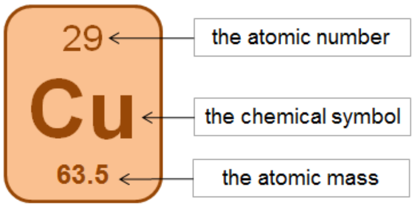
# AQA Trilogy-Chemistry key terms – Atomic Structure and the periodic table

All substances are made up of **atoms**, which are the smallest indivisible particles that make up matter.

Atoms are made of **protons**, **neutrons** and **electrons**. Protons and neutrons are found in the nucleus and electrons are found in energy levels (i.e. shells) around the nucleus.




An atom's atomic and mass number can be found on the Periodic Table. _Use the key on your data sheet_ to determine which number is which!

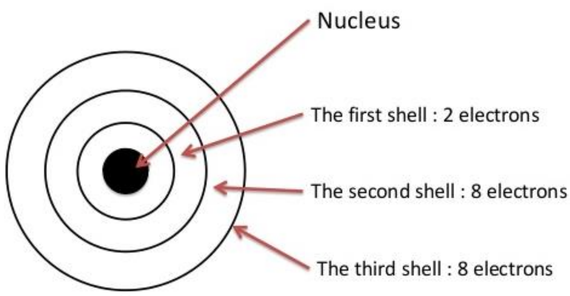



| Subatomic particle | Relative charge | Relative mass | Helps to define… |
|---|---|---|---|
| Proton | +1 | 1 | Atoms |
| Neutron | 0 | 1 | Isotopes |
| Electron | -1 | negligible | Ions |

- **Elements** are pure substances that contain only one type of atom. Atoms of the same element have the same *number of protons* in its nucleus (i.e. their *atomic number*).
- **Atoms** have _no charge_ because each contains an equal number of protons and electrons.
  - ✓ Are very small, about 0.1 nanometer (1 x $10^{-10}$ m)
  - ✓ Almost all the mass of an atom is in the nucleus
- **Isotopes** are atoms that have the same number of protons (ie. atomic number), but a different number of neutrons (i.e. mass number). An atom's **mass number** = number of protons + number of neutrons.
  - ✓ Have similar chemical properties (similar reactions)
  - ✓ Have different physical properties (i.e. density)
- **Ions** are _charged_ particles that result due to there being more or less electrons to protons.
  - ✓ _Negative_ ions are formed when electrons are _gained_
  - ✓ _Positive_ ions are formed when electrons are _lost_

Electrons in an atom are arranged in energy levels (aka: shells). The lowest energy level must be filled first.

The number of electrons in the outermost shell of an atom determines how it reacts.




**The Periodic Table**

The periodic table developed as chemists tried to classify the elements:

- without knowing much about atoms
- with some chemical compounds mistakenly thought to be elements
- without knowing a complete list of the elements.

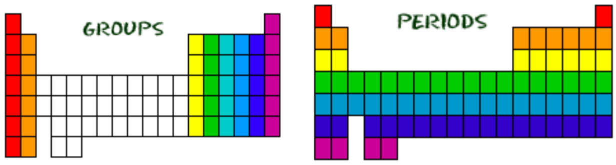
The Periodic Table is so called because of the regularly repeating patterns in the properties of the elements.

| John Dalton (1808) | John Newlands (1864) | Dimitri Mendeleev (1869) |
|---|---|---|
| • Suggested ordering elements by atomic **weight**…<br>• …this led to incomplete versions or placed elements in inappropriate groups.<br>• _Isotopes_ helped to explain why using atomic weights to | • introduced **'law of octaves'** when he noticed that the properties of every eighth element seemed similar.<br>• didn't take into account that new elements were being discovered and his pattern broke down after calcium… | • ordered elements by atomic **number** _and_ their properties, but…<br>• …crucially **left gaps** for unknown elements, which when discovered matched his predictions.<br>• _This table was accepted_ by the scientific community. |

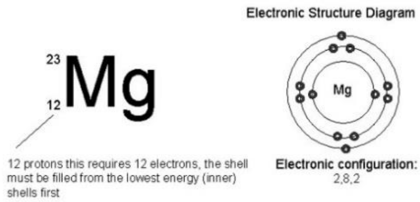
order elements didn't work.

- …because of this, his table was not accepted.

The Periodic Table lists all the chemical elements, organised into **groups** (columns) and **periods** (rows).




Elements found in a **group** have the same number of **electrons in the outermost shell**. Elements found in a **period** have the same number of **shells** of electrons.




Elements are arranged in order of atomic (proton) number.

Metals are found on the left and bottom of the Periodic Table, while non-metals are found on the right and top.

The number of electrons in the outermost shell (highest energy level) of an atom determines its chemical properties.

- Metals tend to lose electrons, forming positive ions.
- Non-metals tend to gain electrons, forming negative ions.
- Noble gases (elements in Group 0) have unreactive because of their very stable electron arrangements (i.e. they have a **full** outer shell of electrons).

You can explain trends in reactivity in terms of the attraction between electrons in the outermost shell and the nucleus.

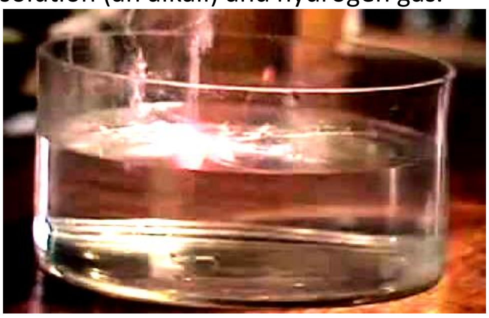

The electrostatic attraction between outer electrons and the nucleus depends on:

- The distance between the outermost electrons and the nucleus (i.e. **atomic radius**). The larger the distance, the weaker the attraction.
- The number of occupied inner shells (energy levels) of electrons. These create a '**shielding'** effect, weakening the electrostatic attraction.
- The number of protons in the nucleus (i.e. **nuclear charge** or 'size' of positive charge)

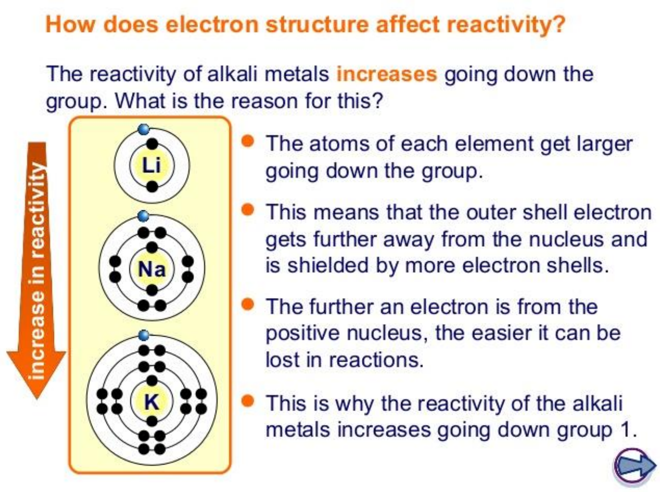
When going down a group, atomic radius and shielding have a larger effect than nuclear charge.
*Larger* atoms tend to *lose* *electrons more easily* and *smaller* atoms tend to *gain electrons more easily*.

**Group 1: Alkali Metals**

- Melting point / boiling point decrease down the group.
- All react with water to produce a metal hydroxide solution (an alkali) and hydrogen gas.

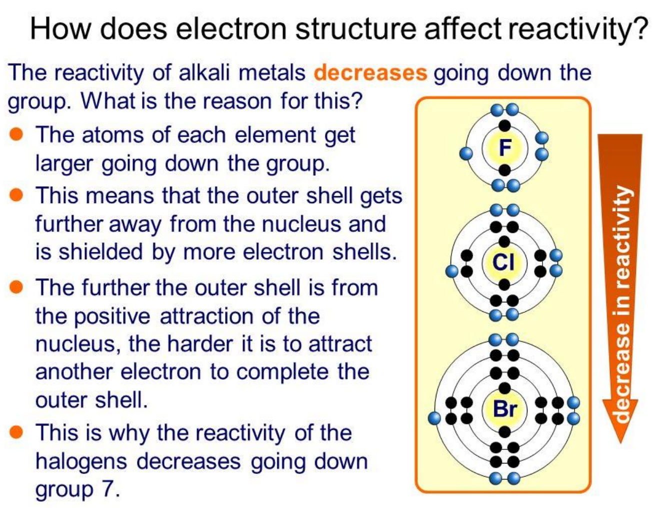


- Have one electron in their outermost shell, so…
- …easily lose one electron to form 1+ ions and make ionic compounds…
- …These compounds are usually white and dissolve in water, showing a colourless solution (i.e. NaCl, table salt, in water)



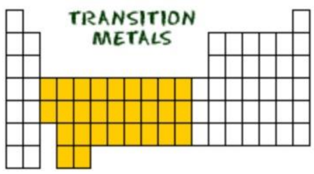
## Group 7: Halogens

- Melting points / boiling points increase down the group.
- Are poor conductors of heat and electricity
- Are all toxic and have coloured vapours
- Exist as diatomic (ie. 2-atom) molecules, e.g. $F_2$, $Cl_2$, etc.
- Have 7 electrons in their outermost shell, so…
- …can gain one electron to form 1- ions and make ionic compounds with metals
- Can also form covalent compounds by sharing electrons with other non-metals
- A more reactive halogen can displace a less eactive halogen from a solution of one of its salts.




## Transition Metals (TRIPLE CHEMISTRY only)

Transition metals are found in the central block of the Periodic Table between Groups 2 and 3.




Many transition metals form coloured compounds, e.g.

| Transition metal ion | Coloured compound |
|---|---|
| $Cu^{2+}$ | blue |
| $Ni^{2+}$ | pale green |
| $Cr^{3+}$ | dark green |
| $Mn^{2+}$ | pale pink |

Have properties typical of metals:

- Good conductors of electricity and heat
- Hard and strong
- Have high densities
- High melting point (except mercury, Hg, which exists as a liquid at room temperature)

**HOWEVER**, compared to Group 1 metals, transition metals are: stronger, harder, have higher melting points and densities, and are less reactive.

Transitions metals **do not** react vigorously with oxygen or water.

Transition metals can form more than one ion. This is why the names of transition metal compounds usually include Roman numerals (i.e. copper(II) sulphate or iron (III) chloride). The Roman numeral indicates what ion the transition metal has formed.

Transition metals and their compounds are important industrial catalysts. For instance, iron is used as a catalyst in the Haber Process to produce ammonia.