

Development of the Periodic Table

We haven't always known as much about chemistry as we do now. Early chemists looked to try and understand patterns in the elements' properties to get a bit of understanding.

In the **Early 1800s** Elements Were **Arranged By Atomic Mass**

Until quite recently, there were two obvious ways to categorise elements:

- 1) Their physical and chemical properties 2) Their relative atomic mass

- 1) Remember, scientists had no idea of atomic structure or of protons, neutrons or electrons, so there was no such thing as atomic number to them.
(It was only in the 20th century after protons and electrons were discovered that it was realised the elements were best arranged in order of atomic number.)
- 2) Back then, the only thing they could measure was relative atomic mass, and so the known elements were arranged in order of atomic mass. When this was done, a periodic pattern was noticed in the properties of the elements. This is where the name 'periodic table' comes from...
- 3) Early periodic tables were not complete and some elements were placed in the wrong group. This is because elements were placed in the order of relative atomic mass and did not take into account their properties.

Dmitri Mendeleev Left Gaps and Predicted New Elements

- 1) In 1869, Dmitri Mendeleev overcame some of the problems of early periodic tables by taking 50 known elements and arranging them into his Table of Elements — with various gaps as shown.

Mendeleev's Table of the Elements

H																
Li	Be											B	C	N	O	F
Na	Mg											Al	Si	P	S	Cl
K	Ca *	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn *	*	As	Se	Br		
Rb	Sr	Y	Zr	Nb	Mo *	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	
Cs	Ba *	*	Ta	W	*	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi			

- 2) Mendeleev put the elements mainly in order of atomic mass but did switch that order if the properties meant it should be changed. An example of this can be seen with Te and I — iodine actually has a smaller relative atomic mass but is placed after tellurium as it has similar properties to the elements in that group.

- 3) Gaps were left in the table to make sure that elements with similar properties stayed in the same groups. Some of these gaps indicated the existence of undiscovered elements and allowed Mendeleev to predict what their properties might be. When they were found and they fitted the pattern it helped confirm Mendeleev's ideas. For example, Mendeleev made really good predictions about the chemical and physical properties of an element he called ekasilicon, which we know today as germanium.

The discovery of isotopes (see page 18) in the early 20th century confirmed that Mendeleev was correct to not place elements in a strict order of atomic mass but to also take account of their properties. Isotopes of the same element have different atomic masses but have the same chemical properties so occupy the same position on the periodic table.

By leaving gaps in the table Mendeleev had the right idea

Make sure you can describe what Mendeleev did to overcome problems with early periodic tables.

The Modern Periodic Table

Chemists were getting pretty close to producing something that you might recognise as a periodic table. The big breakthrough came when the structure of the atom was understood a bit better.

The Periodic Table Helps you to See Patterns in Properties

- 1) There are 100ish elements, which all materials are made of.
- 2) In the periodic table the elements are laid out in order of increasing atomic (proton) number. Arranging the elements like this means there are repeating patterns in the properties of the elements (the properties are said to occur periodically, hence the name periodic table).
- 3) If it wasn't for the periodic table organising everything, you'd have a hard job remembering all those properties.
- 4) It's a handy tool for working out which elements are metals and which are non-metals. Metals are found to the left and non-metals to the right.

relative atomic mass

atomic number

Group 1 Group 2

Group 3 Group 4 Group 5 Group 6 Group 7

Group 8 Group 9 Group 10 Group 11 Group 12 Group 13 Group 14 Group 15 Group 16 Group 17 Group 18

alkali metals transition metals halogens noble gases (see page 39-40) (see page 38) (see page 41-42) (see page 43)

metals non-metals

- 5) Elements with similar properties form columns.
- 6) These vertical columns are called groups.
- 7) The group number tells you how many electrons there are in the outer shell. For example, Group 1 elements all have one electron in their outer shell and Group 7 all have seven electrons in their outer shell. The exception to the rule is Group 0, for example helium has two electrons in its outer shell. This is useful as the way atoms react depends upon the number of electrons in their outer shell. So all elements in the same group are likely to react in a similar way.
- 8) If you know the properties of one element, you can predict properties of other elements in that group — and in the exam, you might be asked to do this. For example the Group 1 elements are Li, Na, K, Rb, Cs and Fr. They're all metals and they react in a similar way (see pages 34-40).
- 9) You can also make predictions about trends in reactivity. E.g. in Group 1, the elements react more vigorously as you go down the group. And in Group 7, reactivity decreases as you go down the group.
- 10) The rows are called periods. Each new period represents another full shell of electrons.



The modern periodic table is vital for understanding chemistry

This is a good example of how science progresses. A scientist has a basically good (though incomplete) hypothesis (see page 1), other scientists question it and bring more evidence to the table. The hypothesis may be modified or even scrapped to take account of available evidence. Only when all of the available evidence supports a hypothesis will it be accepted.

Warm-Up & Exam Questions

The last few pages have been tough with lots of information to learn. Luckily, here are some questions to get your head around to help you test your understanding.

Warm-Up Questions

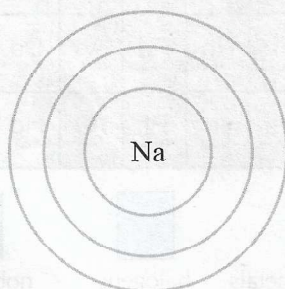
- Describe J J Thomson's model of the atom.
- How many electrons can be held in the following:
 - the first shell of an atom
 - the second shell of an atom
- What discovery supported Mendeleev's decision not to place elements in order of relative atomic mass?

Exam Questions

- 1 Sodium has an atomic number of 11.

Grade
4-6

- 1.1 Complete the dot and cross diagram to show the electron configuration of sodium.



[1 mark]

- 1.2 Sodium is in Group 1.

Name another element that would have the same number of outer shell electrons.

[1 mark]

- 1.3 How many electrons does sodium need to lose so that it has a full outer shell?

[1 mark]

- 2 The periodic table contains all the known elements arranged in order.

Grade
6-7

- 2.1 How were elements generally ordered in early periodic tables?

[1 mark]

- 2.2 How are the elements arranged in the modern periodic table?

[1 mark]

- 2.3 If two elements are in the same group of the periodic table what will they have in common?
Explain your answer.

[1 mark]

- 3* Describe how the theory of atomic structure has changed throughout history.

Grade
7-9

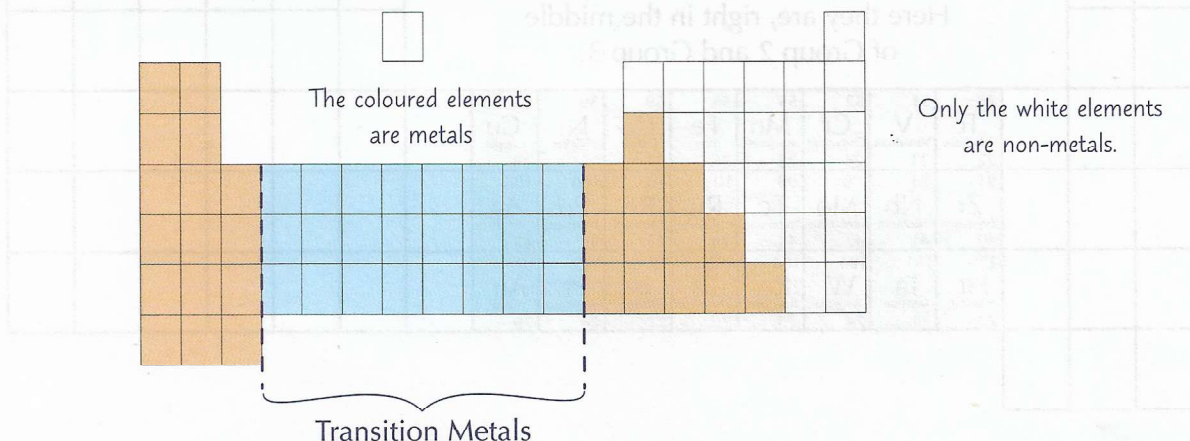
[6 marks]

Metals and Non-Metals

Metals are used for all sorts of things so they're really important in modern life.

Most Elements are Metals

- 1) Metals are elements which can form positive ions when they react.
- 2) They're towards the bottom and to the left of the periodic table.
- 3) Most elements in the periodic table are metals.
- 4) Non-metals are at the far right and top of the periodic table.
- 5) Non-metals don't generally form positive ions when they react.



The Electronic Structure of Atoms Affects How They Will React

- 1) Atoms generally react to form a full outer shell. They do this via losing, gaining or sharing electrons.
- 2) Metals to the left of the periodic table don't have many electrons to remove and metals towards the bottom of the periodic table have outer electrons which are a long way from the nucleus so feel a weaker attraction. Both these effects mean that not much energy is needed to remove the electrons so it's feasible for the elements to react to form positive ions with a full outer shell.
- 3) For non-metals, forming positive ions is much more difficult. This is because they are either to the right of the periodic table — where they have lots of electrons to remove to get a full outer shell, or towards the top — where the outer electrons are close to the nucleus so feel a strong attraction. It's far more feasible for them to either share or gain electrons to get a full outer shell.

Metals and Non-Metals Have Different Physical Properties

- 1) All metals have metallic bonding which causes them to have similar basic physical properties.

- They're strong (hard to break), but can be bent or hammered into different shapes (malleable).
- They're great at conducting heat and electricity.
- They have high boiling and melting points.

- 2) As non-metals don't have metallic bonding, they don't tend to exhibit the same properties as metals. They tend to be dull looking, more brittle, aren't always solids at room temperature, don't generally conduct electricity and often have a lower density.

Non-metals form a variety of different structures so have a wide range of chemical properties.

Metals have quite different properties from non-metals

So, all metals can conduct electricity and heat and be bent into shape, whereas mainly non-metals can't.

Transition Metals

Transition metals make up the big clump of elements in the middle of the periodic table.

Transition Metals can be Found Between Group 2 and Group 3

- 1) Transition metals are in the centre of the periodic table.

Group 1	Group 2											Group 3	Group 4	Group 5	Group 6	Group 7	Group 0
			18 Ti Titanium 22	21 V Vanadium 23	24 Cr Chromium 24	25 Mn Manganese 25	26 Fe Iron 26	27 Co Cobalt 27	28 Ni Nickel 28	29 Cu Copper 29							
			40 Zr Zirconium 40	41 Nb Niobium 41	42 Mo Molybdenum 42	43 Tc Technetium 43	44 Ru Ruthenium 44	45 Rh Rhodium 45	46 Pd Palladium 46	47 Ag Silver 47							
			72 Hf Hafnium 72	73 Ta Tantalum 73	74 W Tungsten 74	75 Re Rhenium 75	76 Os Osmium 76	77 Ir Iridium 77	78 Pt Platinum 78	79 Au Gold 79							

Here they are, right in the middle
of Group 2 and Group 3.

- 2) Transition metals are typical metals, and have the properties you would expect of a 'proper' metal — they're good conductors of heat and electricity, and they're very dense, strong and shiny.
- 3) Transition metals also have some pretty special properties...

Transition metals can have more than one ion.

For example, copper forms Cu^+ and Cu^{2+} ions. Cobalt forms Co^{2+} and Co^{3+} ions.

Transition metal ions are often coloured, and so compounds that contain them are colourful.

For example, potassium chromate(VI), which is yellow, potassium(VII) manganate is purple.

Transition metal compounds often make good catalysts (things that speed up the rate of a reaction — see p.119).

For example, a nickel based catalyst is used in the hydrogenation of alkenes (p.139), and an iron catalyst is used in the Haber process for making ammonia (p.188).

Transition metals are pretty useful elements...

Transition metals have similar properties to 'typical metals' but also have a few more unusual characteristics. These include having multiple ions, coloured compounds, and the ability to act as catalysts.

Group 1 Elements

Group 1 elements are known as the alkali metals — these are silvery solids that have to be stored in oil (and handled with forceps) as they're very reactive.

The Group 1 Elements are Reactive, Soft Metals

- 1) The alkali metals are lithium, sodium, potassium, rubidium, caesium and francium.
- 2) The alkali metals are all soft and have low density. The first three in the group are less dense than water.
- 3) They all have one electron in their outer shell. This makes them very reactive and gives them similar properties.

Trends of Alkali Metals:

The trends for the alkali metals as you go down Group 1 include:

- 1) Increasing reactivity — the outer electron is more easily lost as the attraction between the nucleus and electron decreases, because the electron is further away from the nucleus the further down the group you go.
- 2) Lower melting and boiling points.
- 3) Higher relative atomic mass.

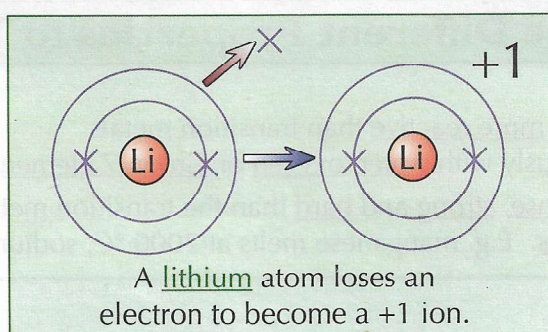
Group 1	Group 2			
7 Li Lithium				
11 Na Sodium				
19 K Potassium				
37 Rb Rubidium				
55 Cs Caesium				
87 Fr Francium				

Alkali Metals Form Ionic Compounds with Non-Metals

- 1) The Group 1 elements don't need much energy to lose their one outer electron to form a full outer shell, so they readily form 1+ ions.
- 2) It's so easy for them to lose their outer electron that they only ever react to form ionic compounds. These compounds are usually white solids that dissolve in water to form colourless solutions.

There's more on ionic compounds on page 50.

Details of the reactions of Group 1 metals with non-metals are on the next page.



The properties of Group 1 metals change as you go down the group

In the exam you might be given a trend and then asked to predict properties of other Group 1 metals. For example, the reactivity of Group 1 metals increases as you go down the group, so you know that potassium will react more vigorously than sodium.

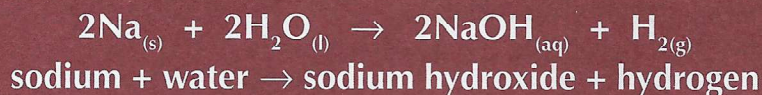
Group 1 Elements

You met Group 1 metals on the previous page, so now it's time to learn about some of their reactions...

Reactions of Alkali Metals with Non-Metals:

Reaction with water

- 1) When Group 1 metals are put in water, they react vigorously to produce hydrogen gas and metal hydroxides — salts that dissolve in water to produce alkaline solutions.

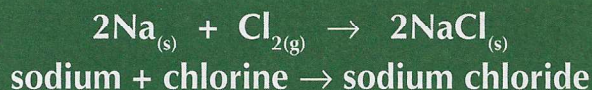


All the Group 1 metals react with water in a similar way.

- 2) The more reactive (lower down in the group) an alkali metal is, the more violent the reaction.
- 3) The amount of energy given out by the reaction increases down the group — the reaction with potassium releases enough energy to ignite hydrogen.

Reaction with chlorine

- 1) Group 1 metals react vigorously when heated in chlorine gas to form white metal chloride salts.



- 2) As you go down the group, reactivity increases so the reaction with chlorine gets more vigorous.

Reaction with oxygen

The Group 1 metals can react with oxygen to form a metal oxide. Different types of oxide will form depending on the Group 1 metal:

- Lithium reacts to form lithium oxide (Li_2O).
- Sodium reacts to form a mixture of sodium oxide (Na_2O) and sodium peroxide (Na_2O_2).
- Potassium reacts to form a mixture of potassium peroxide (K_2O_2) and potassium superoxide (KO_2).

The reactions with oxygen are why Group 1 metals tarnish in the air — the metal reacts with oxygen in the air to form a dull metal oxide layer.

Group 1 Metals Have Different Properties to Transition Metals

- 1) Group 1 metals are much more reactive than transition metals — they react more vigorously with water, oxygen or Group 7 elements for example.
- 2) They're also much less dense, strong and hard than the transition metals, and have much lower melting points. E.g. manganese melts at 2000 °C, sodium melts at 98 °C.



So now you know why they are called alkali metals...

...because they react with water to give alkaline solutions. Make sure you also know how they react with chlorine and oxygen. Memorise a couple of example equations for these reactions.