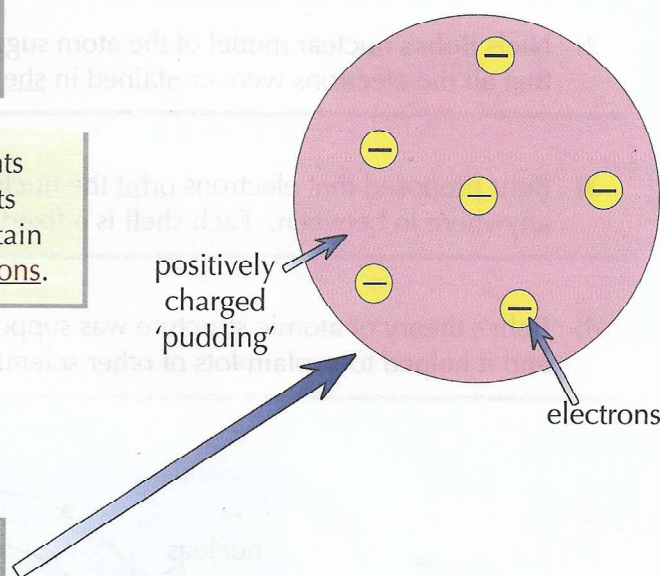


The History of the Atom

You may have thought you were done with the atom after page 16. Unfortunately, you don't get away that easily. The next couple of pages are all about how scientists came to understand the atom as we do today.

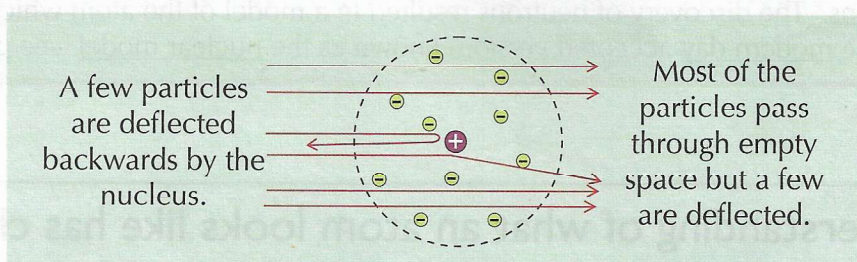
The Theory of Atomic Structure Has Changed Over Time

- 1) At the start of the 19th century John Dalton described atoms as solid spheres, and said that different spheres made up the different elements.
- 2) In 1897 JJ Thomson concluded from his experiments that atoms weren't solid spheres. His measurements of charge and mass showed that an atom must contain even smaller, negatively charged particles — electrons.
- 3) The 'solid sphere' idea of atomic structure had to be changed. The new theory was known as the 'plum pudding model'.
- 4) The plum pudding model showed the atom as a ball of positive charge with electrons stuck in it.



Rutherford Showed that the Plum Pudding Model Was Wrong

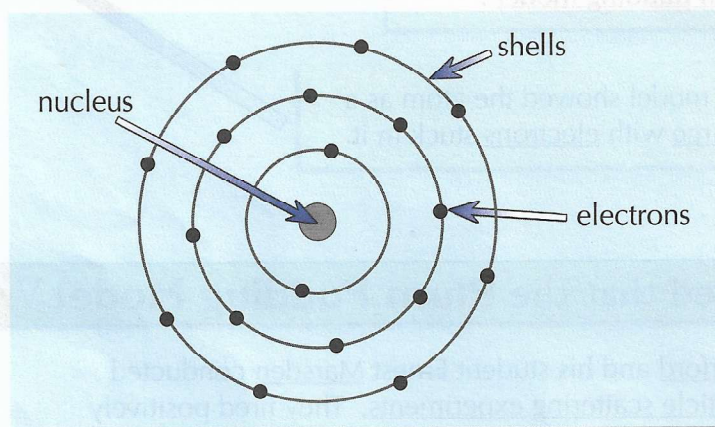
- 1) In 1909 Ernest Rutherford and his student Ernest Marsden conducted the famous alpha particle scattering experiments. They fired positively charged alpha particles at an extremely thin sheet of gold.
- 2) From the plum pudding model, they were expecting the particles to pass straight through the sheet or be slightly deflected at most. This was because the positive charge of each atom was thought to be very spread out through the 'pudding' of the atom. But, whilst most of the particles did go straight through the gold sheet, some were deflected more than expected, and a small number were deflected backwards. So the plum pudding model couldn't be right.
- 3) Rutherford came up with an idea to explain this new evidence — the nuclear model of the atom. In this, there's a tiny, positively charged nucleus at the centre, where most of the mass is concentrated. A 'cloud' of negative electrons surrounds this nucleus — so most of the atom is empty space. When alpha particles came near the concentrated, positive charge of the nucleus, they were deflected. If they were fired directly at the nucleus, they were deflected backwards. Otherwise, they passed through the empty space.



The History of the Atom

Bohr's Nuclear Model Explains a Lot

- 1) Scientists realised that electrons in a 'cloud' around the nucleus of an atom, as Rutherford described, would be attracted to the nucleus, causing the atom to collapse.
- 2) Niels Bohr's nuclear model of the atom suggested that all the electrons were contained in shells.
- 3) Bohr proposed that electrons orbit the nucleus in fixed shells and aren't anywhere in between. Each shell is a fixed distance from the nucleus.
- 4) Bohr's theory of atomic structure was supported by many experiments and it helped to explain lots of other scientists' observations at the time.



Further Experiments Showed the Existence of Protons

- 1) Further experimentation by Rutherford and others gave the conclusion that the nucleus could be divided into smaller particles, each of which has the same charge as a hydrogen nucleus.
- 2) About 20 years after scientists had accepted that atoms have nuclei, James Chadwick carried out an experiment which provided evidence for neutral particles in the nucleus. These became known as neutrons. The discovery of neutrons resulted in a model of the atom which was pretty close to the modern day accepted version, known as the nuclear model (see page 16).



Our understanding of what an atom looks like has changed

Our understanding of the atom has gone through many stages thanks to other people's work being built upon with new evidence and new predictions made. A fine example of the scientific method.

Electronic Structure

The fact that electrons occupy 'shells' around the nucleus is what causes the whole of chemistry. Remember that, and watch how it applies to each bit of it.

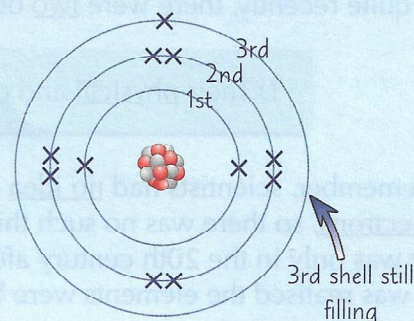
Electron Shell Rules:

- 1) Electrons always occupy shells (sometimes called energy levels).
- 2) The lowest energy levels are always filled first — these are the ones closest to the nucleus.
- 3) Only a certain number of electrons are allowed in each shell:

1st shell: 2 2nd shell: 8 3rd shell: 8

- 4) Atoms are much happier when they have full electron shells — like the noble gases in Group 0.
- 5) In most atoms, the outer shell is not full and this makes the atom want to react to fill it.

Electron configurations can be shown as diagrams like this...



...or as numbers like this: 2, 8, 1

Both of the configurations above are for sodium.

Follow the Rules to Work Out Electronic Structures

You can easily work out the electronic structures for the first 20 elements of the periodic table (things get a bit more complicated after that).

EXAMPLE:

What is the electronic structure of nitrogen?

- 1) Nitrogen's atomic number is 7. This means it has 7 protons... so it must have 7 electrons.
- 2) Follow the 'Electron Shell Rules' above. The first shell can only take 2 electrons and the second shell can take a maximum of 8 electrons.

So the electronic structure for nitrogen must be 2, 5.

EXAMPLE:

What is the electronic structure of magnesium?

- 1) Magnesium's atomic number is 12. This means it has 12 protons... so it must have 12 electrons.
- 2) Follow the 'Electron Shell Rules' above. The first shell can only take 2 electrons and the second shell can take a maximum of 8 electrons, so the third shell must also be partially filled.

So the electronic structure for magnesium must be 2, 8, 2.

Here are some more examples of electronic structures:

H Hydrogen 1 Proton no. = 1	He Helium 2 Proton no. = 2	Li Lithium 2, 1 Proton no. = 3	C Carbon 2, 4 Proton no. = 6	Ne Neon 2, 8 Proton no. = 10	Ca Calcium 2, 8, 8, 2 Proton no. = 20
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Electron shells — one of the most important ideas in chemistry

It's really important to learn the rules for filling electron shells — make sure you practise.

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 0

alkali metals transition metals halogens noble gases (pink line separates metals and non-metals)

(see page 39-40) (see page 38) (see page 41-42) (see page 43)

- 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.