KEEP IT SIMPLE SCIENCE
PhotoMaster Format
Chemistry Module 1

Properties & Structure of Matter

Topic Outline

Properties & Structure of Matter

- **1. Types of Matter**
  - Elements, compounds & mixtures
  - Separating mixtures
  - % composition
- **2. Atomic Structure & the Elements**
  - Basic atomic structure
  - Relative atomic mass
  - Atomic No. & Mass No.
  - Isotopes & Radioactivity
  - Alpha & Beta decay
  - Properties of radiation
  - Occurrence of radioisotopes
- **3. Periodicity**
  - Layout of Periodic Table
  - Electron orbits & orbitals
  - Periodic patterns in the table
- **4. Chemical Bonding**
  - Formation iof ions
  - Ionic bonding & compounds
  - Valency
  - Lewis formulas
  - Covalent bonding & compounds
  - Electronegativity & polar bonding
- **5. Bonding, Structure & Properties**
  - Ionic & covalent lattices, molecules & metals
  - Bonding & structure determines properties
  - Allotropes

What is this topic about?

To keep it as simple as possible, (K.I.S.S. Principle) this topic covers:

1. TYPES of MATTER

Review of elements, compounds & mixtures. Homogeneous & heterogeneous mixtures. Methods for separating mixtures. % composition.

2. ATOMIC STRUCTURE

Basic atomic structure. Atomic Number & Mass Number. Relative Atomic Mass (R.A.M.) Isotopes & radioactivity. Alpha & beta decay. Occurrence & properties of radiation.

3. PERIODICITY

Structure of the Periodic Table. Electron orbits. The Bohr model. Quantum Theory & orbitals. Periodic patterns: atomic radius, electronegativity, ionisation energy, and more.

4. CHEMICAL BONDING

Formation of ions. Ionic bonding & compounds. Names, formulas & equations. Valency. Lewis formulas. Covalent bonds & compounds. Electronegativity & polar bonding.

5. BONDING, STRUCTURE & PROPERTIES

Ionic & covalent lattices, molecules & metals. How bonding & structure determine properties. Allotropes.

1. Properties of Matter

Every substance is either an element, a compound, or a mixture.
It is essential that you understand clearly how each type of matter is different.

Elements

Pure.
Only one type of atom present.

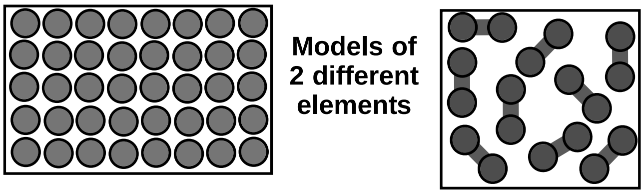
Each element has a unique set of properties.

Listed on the Periodic Table, with its own symbol and Atomic Number.

Cannot be separated into parts by any physical or chemical process.

Examples of Elements

Oxygen, Iron, Copper, Lead, Chlorine



Models of 2 different elements

Compounds

Pure.
Only one type of molecule or ion-pair present.

Each compound has a unique set of properties.

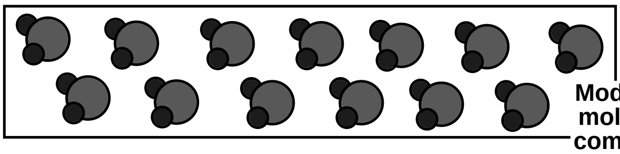
Contains 2 or more elements, chemically bonded together in a fixed ratio.

Cannot be separated into parts by any physical process.

Can be separated into its elements by chemical decomposition.

Examples of Compounds

Water, salt, ethanol, copper sulfate



Model of a molecular compound

In this section we will concentrate on mixtures and how to separate their parts.

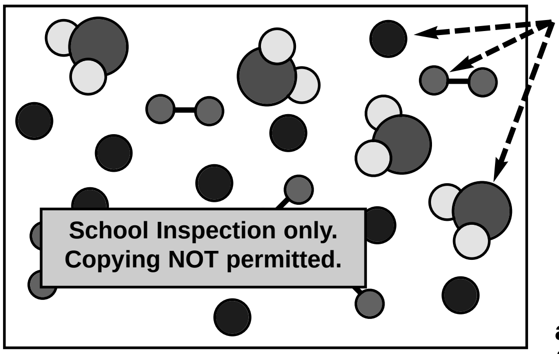
Mixtures

Not pure. (Different particles within.)

Variable composition and properties.

Can be separated into parts by physical processes. (filtering, distilling, etc)

May contain elements and/or compounds within the mix.



Different particles in a mixture

We are Surrounded by Mixtures

The Earth is often thought of as being made up of several "spheres". All are mixtures of compounds and elements.

Lithosphere, Hydrosphere, Atmosphere

The lithosphere is the solid, rocky part of the Earth. Rocks are mixtures of minerals which are, in turn, crystalline compounds. Each type of rock is a different mixure, with different minerals, in varying proportions.

The hydrosphere is the watery part of the Earth, mainly the oceans, rivers and lakes. Most of it is a mixture of water with suspended solids (e.g. dirt), dissolved compounds (especially salt) and dissolved elements such as oxygen and nitrogen.

The atmosphere is a mixture of gases, notably the elements nitrogen & oxygen and compounds such as carbon dioxide.

Homogeneous Mixtures

are those which appear to have a uniform composition throughout, even at microscopic scales.
Homogeneous mixtures are NOT obviously composed of different things mixed together.

Examples include sea water, clean air, or glass (even if coloured). Each appears to be uniform in composition... the different substances within are not visible.

Heterogeneous Mixtures

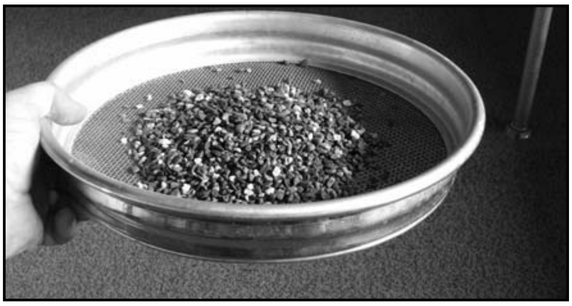
are those which are obviously composed of different substances. Perhaps the classic example is a chunk of concrete in which gravel & sand grains can be clearly seen cemented together. Muddy water can be seen to be heterogeneous when it begins to separate and form a layer of mud within a few minutes of standing.

Separating Mixtures

Since the different particles within a mixture are not all chemically bonded together, and since each has different properties, they can be separated fairly easily by simple physical processes. It is important that you can identify the "Difference in Properties" (D.I.P.) which allows each process to separate the fractions of the mixture.

Solids of Different Grain Size

Imagine a mixture of dry sand and pebbles you have scooped up from a beach. How could the sand be separated from the pebbles?



Using a Sieve

Fine material (sand) falls through the mesh. Coarser pebbles are caught.

D.I.P = grain sizes

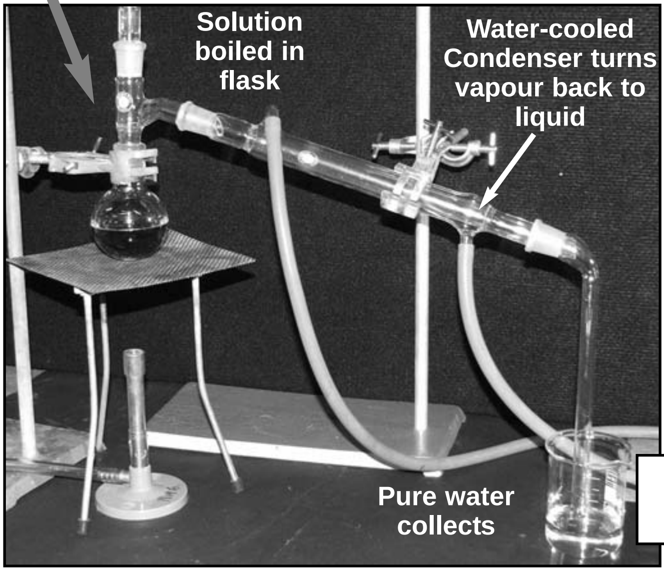
Dissolved Solids in Liquids

When a solid is dissolved in a liquid, such as when salt dissolves in water, the mixture is called a "solution" and filtration will not work to separate the parts.

Later in this course you will learn in detail what happens when solids dissolve. At this point just be aware that in a solution the particles of the dissolved solid ("solute") are similar in size to the molecules of the liquid ("solvent"). If the water molecules can get through the filter paper, the dissolved solute particles will too.

The commonest ways to separate the parts of a solution are:

Evaporation... **to collect the solid solute, and**
Distillation... **to collect the liquid solvent.**

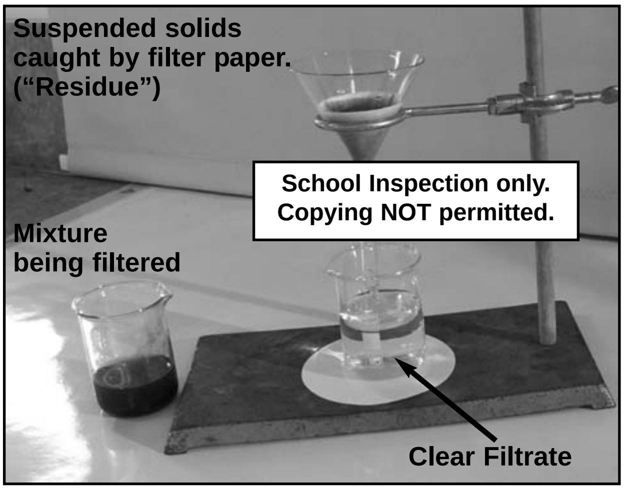


Distillation collects the solvent from a solution.

Solids and Liquids (when NOT a solution)

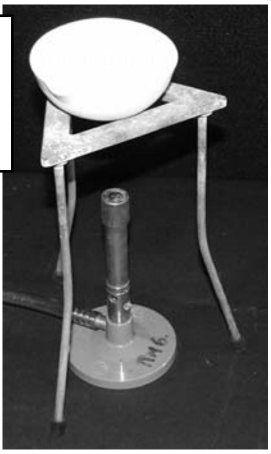
If a solid is suspended in a liquid (such as sand mixed with water) it will often separate by itself if allowed to stand. When a solid settles-out of a suspension like this, it is called sedimentation.

In the laboratory or in industry, separation can often be achieved faster and more efficiently by filtration.



D.I.P = particle size

A filter paper is like a "sieve" of paper fibres, with many small holes. Water molecules can pass through the holes, but the larger particles of the suspended solid are caught.



Simple evaporation procedure. This collects the solute from a solution.

For example, with a salt-water solution, the water boils (and vaporises) at 100^{o}C. The salt however, wouldn't even melt until 770^{o}C and so it stays in the basin or flask.

D.I.P.= different boiling points. (b.p.)

As the water evaporates away the salt solution becomes more and more concentrated, until solid salt crystals begin to separate from the remaining solution. In a distillation, it is time to stop heating before the flask over-heats and breaks!

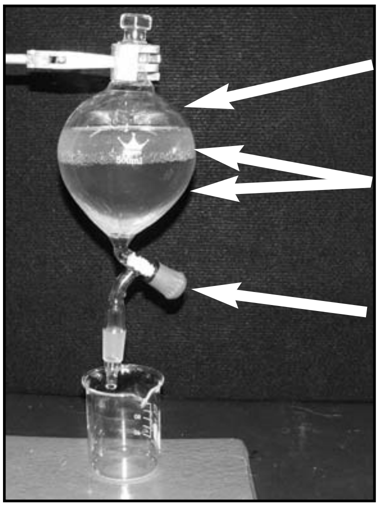
Separating Mixtures **cont.**

Separating Liquid Mixtures

If 2 liquids can mix together and dissolve in each other (like alcohol in water, or oil in petrol) they are said to be "miscible". If 2 liquids will not mix with each other (like oil and water) they are "immiscible".

Separating immiscible liquids can be easily done with a separating funnel.

D.I.P. = immiscible & different density



If the liquids are miscible, separation is more difficult.

If their boiling points are quite different, distillation will work.

D.I.P.= different boiling points.

However, if the b.p.'s are similar, it might be difficult to get total separation into really pure "fractions". For example, when distilling alcohol-water mixtures it is impossible to collect pure alcohol, and in the industrial distillation of (say) wine to make brandy, the distillate is about 40% alcohol, 60% water.

Separating Gas Mixtures

For example, how could you separate air into its different gases? The technique used is called "Fractional Distillation".

D.I.P.= different boiling points.

Basically, air can be turned to liquid, by cooling and compressing it. Then, if allowed to gradually warm up, each different gas "fraction" boils off at its particular b.p., and can be collected separately... pure oxygen, pure argon, etc.

Fractional Distillation is also used to separate crude oil (petroleum) into petrol, kerosene, diesel fuel, etc.

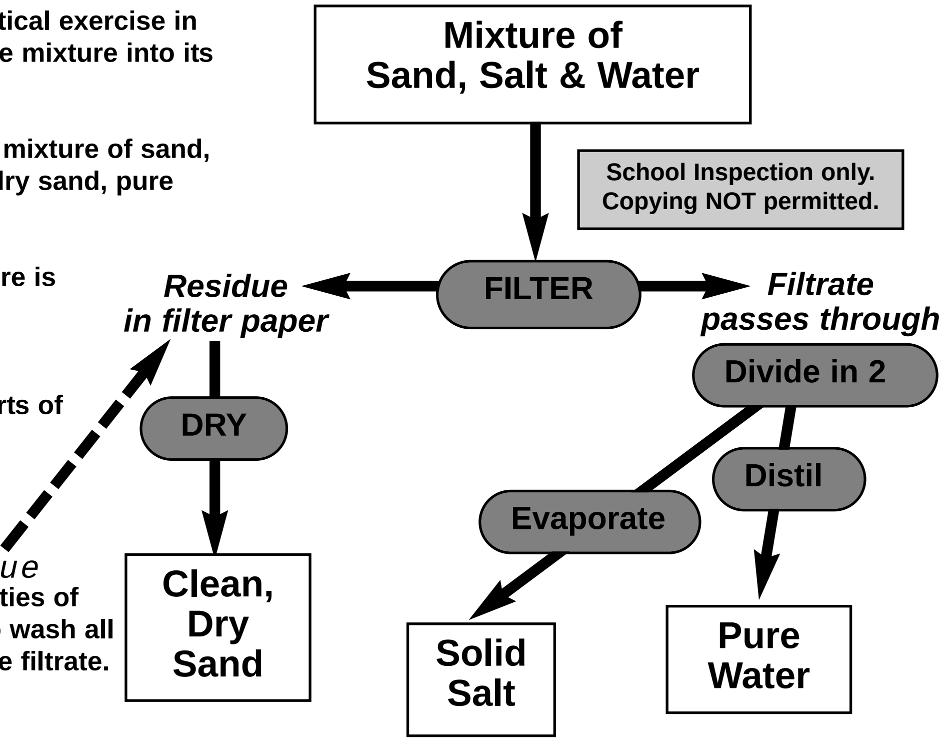
Practical Work: Separating a Mixture

You may do, or have done, a practical exercise in the laboratory to separate a simple mixture into its fractions.

A common task is to begin with a mixture of sand, salt and water and collect clean, dry sand, pure solid salt and pure water.

A flow-chart of a suitable procedure is shown.

You need to be able to interpret, and to construct, similar flow charts of procedures.



A Point of Good Technique

It is important to add small quantities of extra, pure water to the residue to wash all the dissolved salt through with the filtrate.

Gravimetric Analysis: Percentage Composition

To separate the parts of a mixture is one thing, but very often in industry or science it is important to measure the quantities or percentages of each fraction.

Practical Work...

A Simple Gravimetric Analysis

You may have carried out a practical exercise similar to that shown in the flow-chart.

If you were given a solid mixture of <u>sand and salt</u>, how could you determine the percentage of each in the mixture?

Analysis of Results

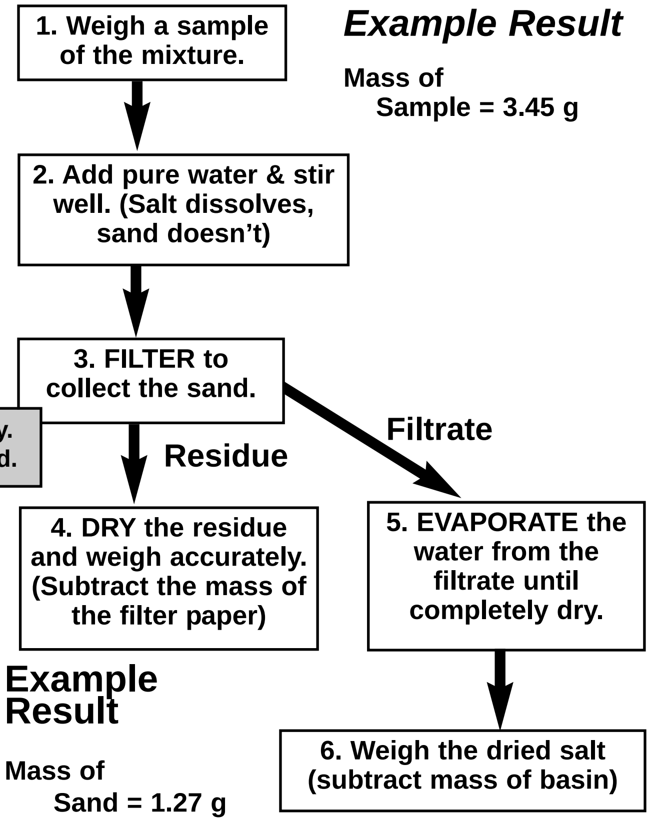
$$\%\ \text{Sand in sample} = \frac{\text{mass of sand}}{\text{mass of sample}} \times 100$$

$$= (1.27 / 3.45) \times 100$$
$$= 36.8\%$$

$$\%\ \text{Salt in sample} = \frac{\text{mass of salt}}{\text{mass of sample}} \times 100$$

$$= (2.08 / 3.45) \times 100$$
$$= 60.3\%$$

Outline of Procedure



1. Weigh a sample of the mixture.

Example Result

Mass of Sample = 3.45 g

2. Add pure water & stir well. (Salt dissolves, sand doesn't)

3. FILTER to collect the sand.

4. DRY the residue and weigh accurately. (Subtract the mass of the filter paper)

Example Result

Mass of Sand = 1.27 g

5. EVAPORATE the water from the filtrate until completely dry.

6. Weigh the dried salt (subtract mass of basin)

Example Result

Mass of Salt = 2.08 g

<u>Notes</u>:

1. The calculated percentages add to only 97.1%, so there must have been some other substance(s) present, OR this represents the experimental error. (Possibly there was moisture in the mixture which doesn't get accounted for. Perhaps you need to dry the mixture sample thoroughly, before starting.)

2. It is appropriate to round-off the final answers to 3 significant figures as shown, since that matches the precision of the experimental measurements.

Examples of where this analysis might be important:

Measuring the amount of a mineral within an ore deposit to decide if it is economically worth mining it.

Measuring quantities of pollutant chemicals in water or air.

To check that manufacturers' claims are correct regarding the chemical composition (e.g. fat or salt content) of food.

Gravimetric Analysis involves separating the parts of a mixture and accurately measuring the masses along the way, so that you can calculate the composition of the mixture.

2. Atomic Structure & the Elements

Every Element has its Own Atoms

You should already be aware that everything is made of tiny “lumps” of matter called “atoms”. Each atom sometimes acts as if it was a solid ball, but we know that each one is actually made up of even smaller particles.

Protons, Electrons & Neutrons

The <u>electrons</u> are whizzing around the central nucleus, like miniature planets around the Sun.

Each electron, and each proton in the nucleus, carries an electrical charge.

The electrons have a lot of energy and move rapidly. They would instantly fly off in all directions except for their electrical attraction to the protons in the nucleus. So, the orbit of an electron is the “balance” between its fast movement and the electrical attraction pulling it towards the nucleus.

<u>Protons</u> in the nucleus repel each other electrically, so why doesn’t the nucleus fly apart?

The protons and neutrons in the nucleus are held together by an even more powerful force called simply the “<u>strong nuclear force</u>”. This force easily overpowers electrical repulsion.

Electrons carry negative charge. Neutrons have NO charge. Protons carry positive charge.

One Type of Atom = A Chemical Element

The atoms of each element are all basically the same as each other, but different to the atoms of another element.

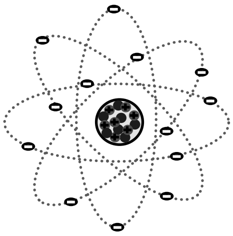
How are the atoms of different elements different?

The atoms of each element have a certain number of protons, electrons and neutrons.

For example, the atoms of aluminium all have:

13 protons
13 electrons
14 neutrons
(not all shown here)

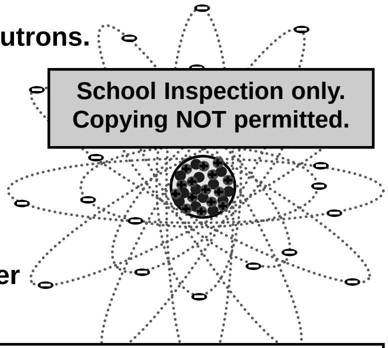
Every atom of aluminium is the same.



Atoms of copper all have:

29 protons
29 electrons
34 neutrons

Every atom of copper is the same.



Most of this material should be revision of known facts.
Skip it if already understood!

Notice that in each case
Number of protons = Number of electrons.
This is true for every element, and defines that element.
Protons = Electrons = “Atomic Number”

The Mass of Atoms

Obviously a single atom has an extremely small mass if you measure it in grams. Instead of using tiny fractions of a gram, we usually consider the <u>relative</u> mass of different atoms. To find relative mass, we simply compare <u>how many particles</u> each atom has.

Mass of Protons, Neutrons & Electrons

It turns out that protons and neutrons have almost the same mass, so we use this as the unit to compare the mass of atoms. This amount of mass is called the “atomic mass unit” or “amu”.
(1 amu is about 1 million billion billionth of a gram)

Electrons are so small ($^{1}/_{1,800}$ amu) that, for all practical purposes, <u>they can be ignored</u> when working with atomic mass. The relative mass of an atom can be found by adding together the number of protons and neutrons in the nucleus.

Particle	Charge	Mass
Proton	+1	1 amu
Electron	-1	$^{1}/_{1,800}$
Neutron	0	1 amu

Atomic Mass Number

Atomic Mass Number = No. of Protons + No. of Neutrons

This must be a whole number for any atom.
(You can’t have half a proton!)

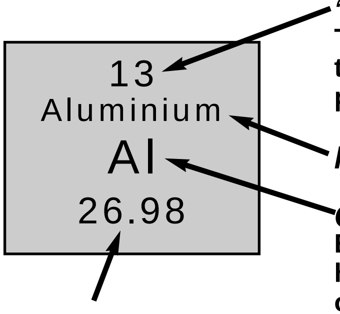
The “Mass Number” for each atom is also known as the “Nucleon Number”.
(“Nucleon” refers to any particle in the nucleus... proton or neutron)

Simplicity Warning!

Be aware that the information above is approximate only. Protons & neutrons are NOT precisely 1 amu each. We are applying the KISS Principle, but you may be required to learn the formal definition of the amu & further details.

Reading the Periodic Table

How to Read the Information for Each Element



“Atomic Number” ***(= protons = electrons)***
The elements are numbered, in order, across and then down the table. This puts the elements in a numerical order, but it also tells you how many electrons & protons in each atom.

Name of the Element

Chemical Symbol
Each element has a short-hand symbol. It is always one capital letter, OR if 2 letters, always a capital followed by a lower case letter.

Shorthand for an Atom

From this, you can work out that each Argon atom contains 18 electrons, 18 protons and 22 neutrons.

$^{40}_{18}\text{Ar}$

Mass No. above
Chemical symbol
Atomic No. below

“Relative Atomic Mass” (RAM) **(Sometimes loosely called the “Atomic Weight”)**
This gives the relative mass of an “average” atom of the element.

Whoa! Wait a minute! That cannot be correct!
The Atomic Mass Number (Nucleon Number) must be a whole number. What’s going on?

The explanation lies in a knowledge of *ISOTOPES*...

Isotopes

You know that the atoms of an element are all the same as each other. **<u>Actually, that’s not quite true</u>!**

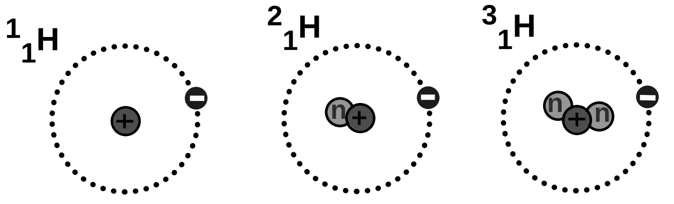
All the atoms of an element have exactly the same
- **number of protons**
- **number of electrons**

} = “Atomic Number” or “Nucleon Number”

It is the number of electrons which gives each atom its chemical properties, and defines it as a particular element. however, the number of neutrons can vary.

For example, hydrogen has the smallest, simplest atoms of all, but there are 3 variations, or “isotopes”.

Example: Isotopes of Hydrogen



0 neutrons **1 neutron** **2 neutrons**

These atoms have the same chemistry, because the electrons are the same ; they are all Hydrogen.

However, their Mass Numbers are different.

> **ISOTOPES are atoms of the same element (same Atomic Number) but with different numbers of neutrons and different ATOMIC MASS NUMBERS. Their chemistry is the same.**

Most elements exist in nature as a mixture of 2 or more isotopes. The “<u>Relative Atomic Mass</u>” shown on the Periodic Table is the <u>average</u> of the mixture of isotopes that occurs on Earth.

Example: Chlorine has 2 main isotopes

$^{35}_{17}\text{Cl}$	$^{37}_{17}\text{Cl}$
17 protons	**17 protons**
17 electrons	**17 electrons**
<u>18 neutrons</u>	**<u>20 neutrons</u>**

On Earth, there is a mixture of these 2 isotopes in such a proportion so that the “average” atomic mass is 35.45. This is the value of R.A.M. shown in the Periodic Table.

Isotopes are commonly described by their individual mass numbers. The isotopes above are called “Chlorine-35” and “Chlorine-37”, or simply Cl-35 and Cl-37.

The isotopes of Hydrogen (at left) are Hydrogen-1, Hydrogen-2 and Hydrogen-3.

Calculating R.A.M. from Isotope Data

If you know the proportions of each isotope in a sample of an element, you can easily calculate the R.A.M.

<u>Example</u>: A sample of chlorine is found to contain 77% Cl-35 and 23% Cl-37.

R.A.M. = (nucleon no. x %) + (nucleon no. x %) +... for each isotope

$$= 35 \times \frac{77}{100} + 37 \times \frac{23}{100} = 35.46$$

The slight variation between this calculated value & the RAM shown in a Periodic Table is due to slight differences in the mix of isotopes in different samples of the element. The official RAM in the Periodic Table is the average of many, many measurements.

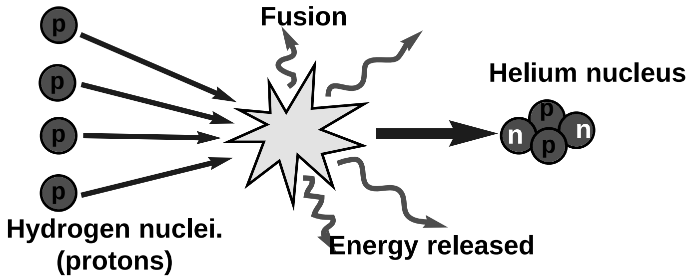
Chemistry is mostly about chemical reactions, and these are controlled by the electrons in atoms. However, since the story of isotopes has arisen, it is appropriate (and a syllabus requirement) to deal with some reactions & transformations which occur in the nucleus of an atom.

Nuclear Reactions

The nucleus of every atom is held together by the “strong nuclear force”. This force is the strongest known. It is far stronger than electrical or magnetic forces, and billions of times stronger than gravity. Certain kinds of changes in the nucleus can release some of this Nuclear Energy.

Nuclear Fusion

is when 2 small atomic nuclei are slammed together so hard that they join and become one. They join to form a larger nucleus and in the process some nuclear energy may be released.

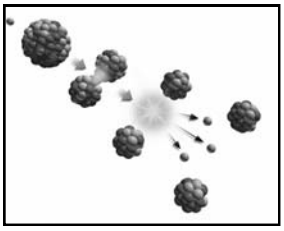


This type of reaction is called “Nuclear Fusion”. It is the process which powers the stars. In a star, hydrogen is fused into helium. Helium can later be fused to form carbon atoms and so on.

In fact, we believe that the Universe was originally made entirely of very small atoms. All the larger atoms have been made by fusion in the stars.

On Earth, the fusion process only occurs in a “Thermonuclear (Hydrogen) Bomb”. We would like to be able to use nuclear fusion for peaceful energy production, but so far we have not figured out how to control the process safely.

Nuclear Fission



Under certain conditions, a very large atomic nucleus (e.g. uranium or plutonium) can break apart into smaller fragments.

A nucleus which splits may release nuclear energy. It also can set off other nuclei, so the result is a “fission chain reaction”.

This is the process in a nuclear reactor used to generate electricity in many countries. It is also the energy source in an “atomic bomb”.

In a nuclear power station the chain reaction is controlled. The energy is used to make steam to drive an electrical generator.

**In a bomb, the chain reaction runs out of control and releases the energy instantly...
a nuclear explosion.**

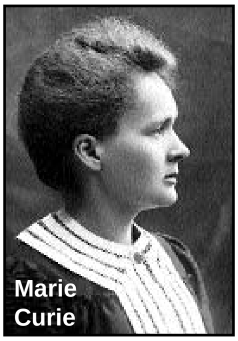
At this stage, the syllabus does NOT require a knowledge of fusion or fission. It is shown here for background information.

The next bit IS required by the syllabus! (Although the history is just for interest)

Radioactivity

In 1896, the French scientist Henri Becquerel accidentally discovered that certain minerals containing uranium were emitting a mysterious, invisible radiation. This was later called “radioactivity”, meaning that the substance was actively emitting radiation.

After Becquerel’s discovery, scientists soon discovered that these radiations were coming from inside the atoms of uranium.



Marie & Pierre Curie were the leaders in this research & after Pierre’s tragic death, Marie continued the work until her death in 1934 from a blood disorder probably caused by her exposure to radiation.

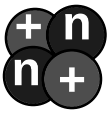
By the 1930’s the research of the Curies, and others, had established:

- **the nature of radioactivity.**
- **that the radiation was coming from the nucleus of atoms.**
- **the occurrence of different isotopes of each element.**
- **that some isotopes are “stable” (do NOT emit radiation), but others are “unstable” which causes them to be “radioisotopes”.**

Quite early on it was discovered that there were, in fact, 3 different radiations. They were quickly labelled alpha (α), beta (β) and gamma (γ) rays. We now know they come from the nucleus of atoms.

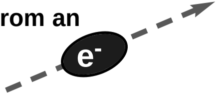
Alpha Radiation

is a stream of particles. An alpha particle is a “chunk” of nucleus, made up of 2 protons and 2 neutrons.



Beta Radiation

is also a stream of particles: this time it is high-speed electrons ejected from an atomic nucleus.



Gamma Radiation

is very high frequency electromagnetic waves, similar to X-rays, but carrying even more energy.

Gamma radiation is often associated with the emission of alpha and beta particles.

Alpha (α) Decay

Every atomic nucleus is held together by the “strong nuclear force”. While this is very powerful, it is also very short-ranged and depends on a certain “balance” of protons & neutrons. If this balance is wrong, or the nucleus is very large, it can be unstable. It may undergo a nuclear reaction to change into a more stable form. The process can involve the emission of particles and radiation... radioactivity.

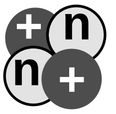
Alpha decay occurs in atoms which have a very large nucleus and are unstable. To achieve greater stability, the nucleus “spits-out” an alpha particle to get rid of excess mass and energy. As it does this, the nucleus turns into a different element. This decay may occur over and over, until the large, unstable atom “decays” into a smaller, stable atom such as lead.

Example

**Uranium is well known as a radioactive substance, and “nuclear fuel” for nuclear reactors and bombs.
Its most common isotope is U-238, meaning it has a mass number of 238. It decays as follows:**

$$^{238}_{92}\mathrm{U} \longrightarrow {}^{234}_{90}\mathrm{Th} + {}^{4}_{2}\mathrm{He}$$

Uranium-238 Thorium-234 Alpha particle



The α-particle consists of 2 protons & 2 neutrons. It is the nucleus of a Helium atom, so it is usually written as $^{4}_{2}\mathrm{He}$

Note that the Mass No. decreases by 4, and the Atomic No. decreases by 2.

Example 2

Radium-226 undergoes alpha decay:

$$^{226}_{88}\mathrm{Ra} \longrightarrow {}^{222}_{86}\mathrm{Rn} + {}^{4}_{2}\mathrm{He}$$

Radium-226 Radon-222 Alpha particle

**Hints: Use the Periodic Table to find Atomic Numbers and identify names and symbols.
Note that the mass numbers AND the atomic numbers ALWAYS BALANCE (add up)
on each side of a “decay equation”.**

Beta (β) Decay

Some atomic nuclei, of any size, have an unstable mix of protons and neutrons. If there is an excess of neutrons, a “nuclear reaction” occurs which converts a neutron into a proton, plus an electron.

$$\text{Neutron} \longrightarrow \text{Proton} + \text{Electron}$$

How can this happen? It seems like magic, but it shows what a strange place the nucleus is. You may study more detail later; for now you must accept that it actually happens.

To understand a “decay equation” for β-decay, you need to know that electrons can be described by the following shorthand.

$$^{0}_{-1}e^{-}$$

Electrons have such little mass that it counts as zero.

To make everything “balance” in a decay equation, the Atomic Number is taken to be -1.

Results of Beta-Decay

**Number of neutrons decreases by 1 and protons increase by 1. (So Atomic Number goes up 1 but Mass Number does not change)
An electron is created in the nucleus, then ejected at high speed. This is the Beta particle...
a high speed electron.**

Examples of b-Decay

Carbon-14 is a well-known radioactive isotope which undergoes beta decay:

$$^{14}_{6}\mathrm{C} \longrightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}e^{-}$$

Carbon-14 Nitrogen-14 Beta particle

In some cases of beta-decay there may be a gamma ray emitted as well, but you cannot predict which ones do, or do not emit gamma rays.

Iodine-131 is a radioactive beta-decayer which also emits a gamma ray:

$$^{131}_{53}\mathrm{I} \longrightarrow {}^{131}_{54}\mathrm{Xe} + {}^{0}_{-1}e^{-} + \gamma$$

Iodine-131 Xenon-131 Beta particle Gamma ray

**Note that once again the Mass Numbers and Atomic Numbers ALWAYS BALANCE across the equation.
(Gamma emission does not affect the numbers)**

> *Note:*
> **There are other decay mechanisms which will not be covered here. If interested, research “positrons” & “electron-capture” to learn more.**

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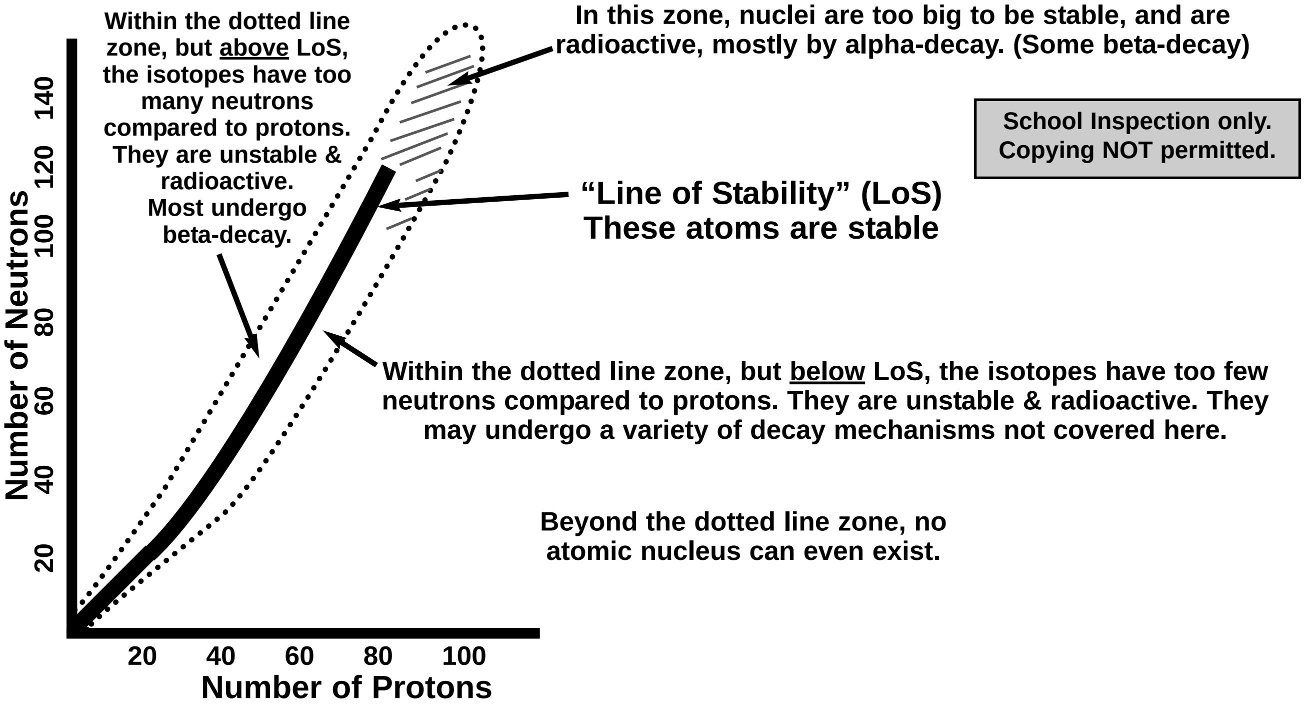
Stable & Unstable Isotopes on the Periodic Table

The syllabus requires that you examine the position of stable & unstable isotopes in the Periodic Table. In one sense, this is a trivial task. Some elements have a dozen or more isotopes, but every isotope of any given element has the same Atomic Number, symbol & element name. Therefore, they all occupy THE SAME position in the table.

However, there are certain patterns. There are very few stable isotopes of really large atoms. Any isotope with atomic number more than about 80 is quite likely to be radioactive, very likely by alpha-decay.

Among smaller atoms, alpha-decay is very unlikely. However, an atom can't exist with just any combination of protons & neutrons in its nucleus. The nucleus can only be stable if the neutron/proton ratio is within certain, very narrow limits. This is best seen if all the known isotopes of all the elements are plotted on a graph, as follows:

Graph of Neutron/Proton Ratio
for all known isotopes (approximate)

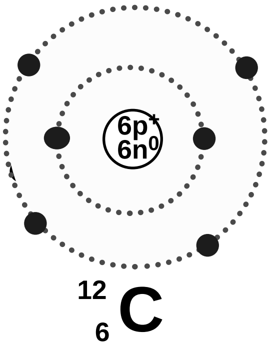


As an example, consider the isotopes of carbon:

Carbon-12, if plotted on the graph (above), lies exactly on the "Line of Stability".

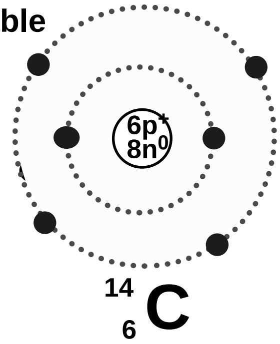
It is stable and NOT radioactive.

In contrast, Carbon-14 lies just above the "Line of Stability", inside the "dotted-line zone".



The C-14 nucleus is unstable because it has too many neutrons, (or not enough protons) in its neutron/proton ratio.
To achieve stability, it undergoes a Radioactive Decay to get rid of some electric charge and some energy.

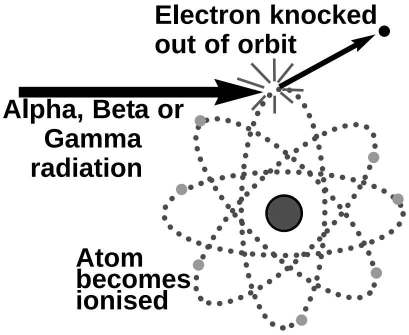
It emits a beta particle and <u>transmutes</u> into a stable nitrogen atom. The decay equation appears on the previous page.



Properties of Alpha, Beta & Gamma Rays

Radiation Causes Ionisation

All 3 radioactive radiations can cause ionisation... they can cause electrons to be knocked out of their orbit around an atom, turning the atom into an ion.



This is why radiation is dangerous to living things. Ionisation of atoms in a living cell can disrupt membranes, cause genetic mutations or alter the cell's DNA so that it becomes cancerous.

The massive ALPHA particle has the highest ionisation ability, BETA is much less ionising and GAMMA less again.

Penetrating Ability

Alpha, Beta and Gamma radiation are quite different in their abilty to penetrate through different substances. You may do <u>Practical Work</u> on this.

Typical Practical Activity is suggested by the diagram.

Typical Findings

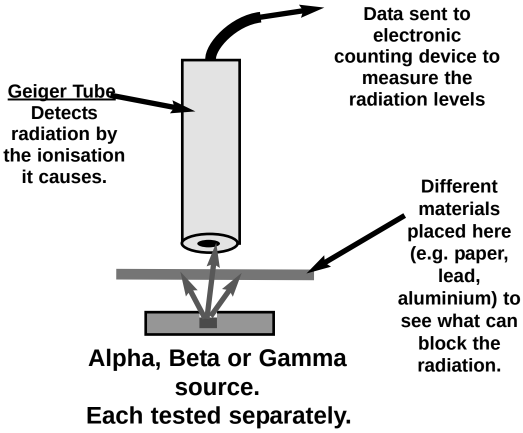
<u>ALPHA particles have low penetrating ability</u>.

A few centimetres in air is as far as they'll get, and a piece of paper will stop 99% of them.

<u>BETA particles penetrate further than alpha</u>,

but rarely go more than 10-20cm in air and most can be stopped by thin metal sheets such as aluminium foil.

<u>GAMMA rays are highly penetrating</u>.
Gamma can travel many metres through air and other substances. To absorb gamma rays, several centimetres of lead or a metre of concrete is just a good start.



Occurrence of Radioisotopes

How common are they? Where are they?

Radio-Isotopes in Nature

Small amounts of radioisotopes occur in rocks & soil, in the air and water and in your food. There is constantly a low-level of "background radiation" around us.

Most of it comes from the remnants and decay products of radioactive atoms which were present when the Earth formed nearly 5 billion years ago. Over the ages, these isotopes decay into stable atoms, so there is less and less of them over time.

A few are constantly produced by natural processes. For example, the well-known radioisotope <u>Carbon-14</u> is constantly produced by nuclear reactions in the upper atmosphere, caused by "cosmic rays" from outer space. The production of C-14 balances its decay rate , so its levels remain fairly constant over time.

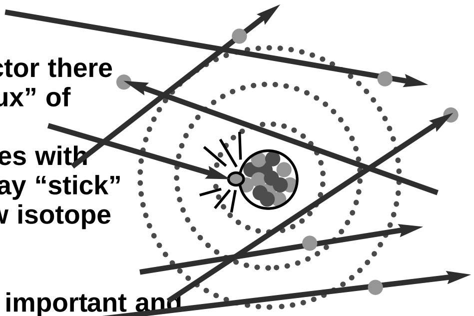
For the last 60-70 years the levels of some radioisotopes in the environment have increased due to human activities. Until banned by treaty, many countries carried out atomic bomb tests. There have also been accidental releases from disasters such as the 2011 explosion at the <u>Fukushima Nuclear Power Station</u> (Japan) following severe damage caused by an earthquake and tsunami.

Artificial Radio-Isotopes

Some radioisotopes are extremely useful for scientific research, medicine and various industries. Some of the most useful do not occur naturally.

To meet these needs, useful isotopes can be made inside a nuclear (fission) reactor. Generally, this is achieved by placing the appropriate "target" atoms inside a nuclear reactor and allowing neutrons to bombard them.

**In a nuclear reactor there is a constant "flux" of neutrons.
When one collides with the nucleus it may "stick" and create a new isotope of that element.**



One of the most important and commonly used radioisotopes produced is Cobalt-60. It is produced when "ordinary", stable Cobalt-59 absorbs a neutron:

$$^{59}_{27}\mathrm{Co} + ^{1}_{0}\mathrm{n} \longrightarrow ^{60}_{27}\mathrm{Co}$$

Most of the world's supply is made in Canada.

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3. Periodicity

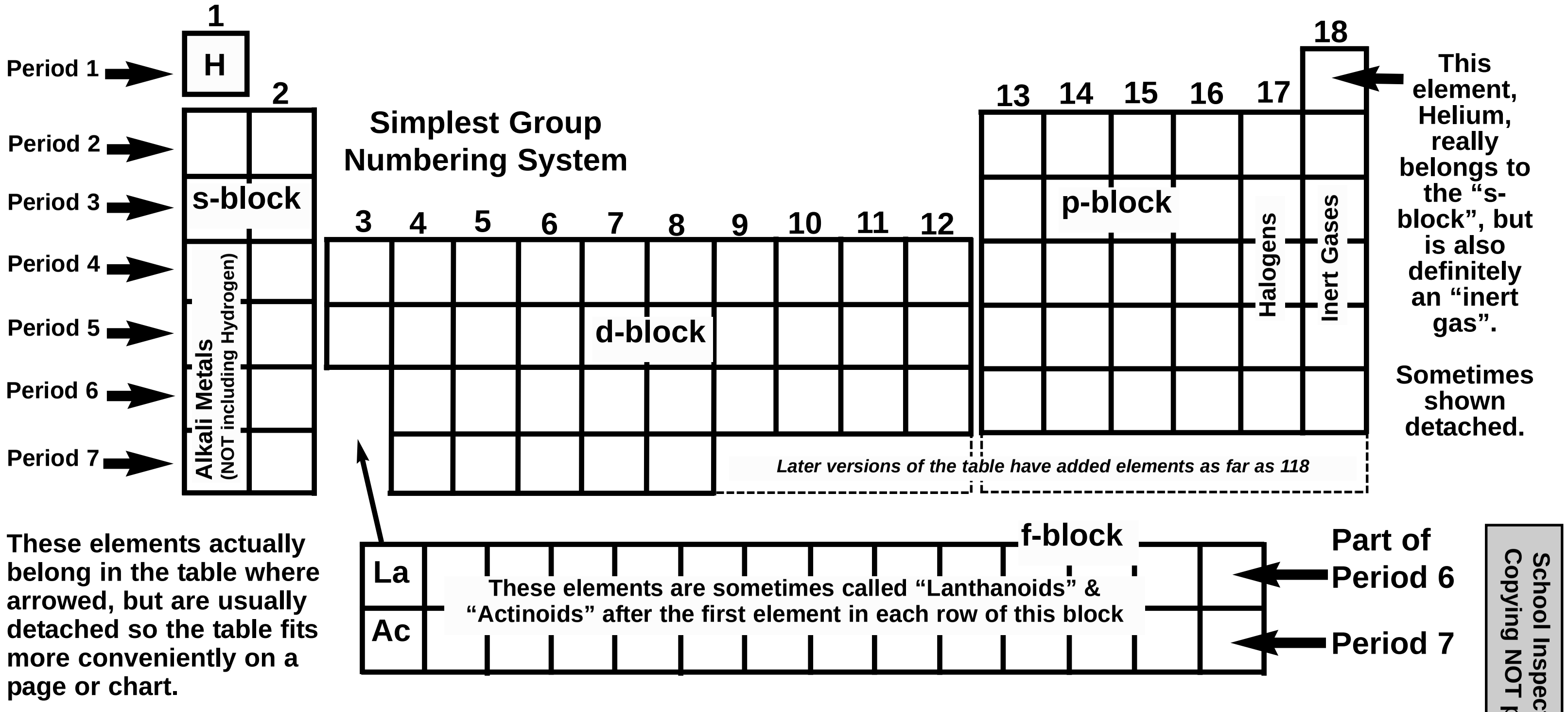
"Periodicity" is a noun which refers to anything which repeats itself or shows recurring patterns. The adjective is "periodic".

Why is the Periodic Table such a weird shape?

Why not put the elements into a simple rectangular box table?
The Periodic Table has its odd shape so that elements that are similar to each other are under each other, or in "groups" and "blocks'. It is called "periodic" because it has patterns that re-occur in a regular pattern. In this section you will learn some of these patterns.

Naming the Rows, Columns & "Blocks"

1. The horizontal rows are called "periods" and are simply numbered as shown in down the left side.
2. The vertical columns are called "groups". There are a variety of ways to number these. The simplest method is shown across the top of the table.
As well as numbering systems, many of the vertical groups have a common name. Some of the best known are shown.
3. The main "blocks" of elements (detached from each other in this diagram) are known by the naming system shown. These block names are connected to electron orbit details and will be explained soon.



This element, Helium, really belongs to the "s-block", but is also definitely an "inert gas".

Sometimes shown detached.

These elements actually belong in the table where arrowed, but are usually detached so the table fits more conveniently on a page or chart.

Classifying the Elements

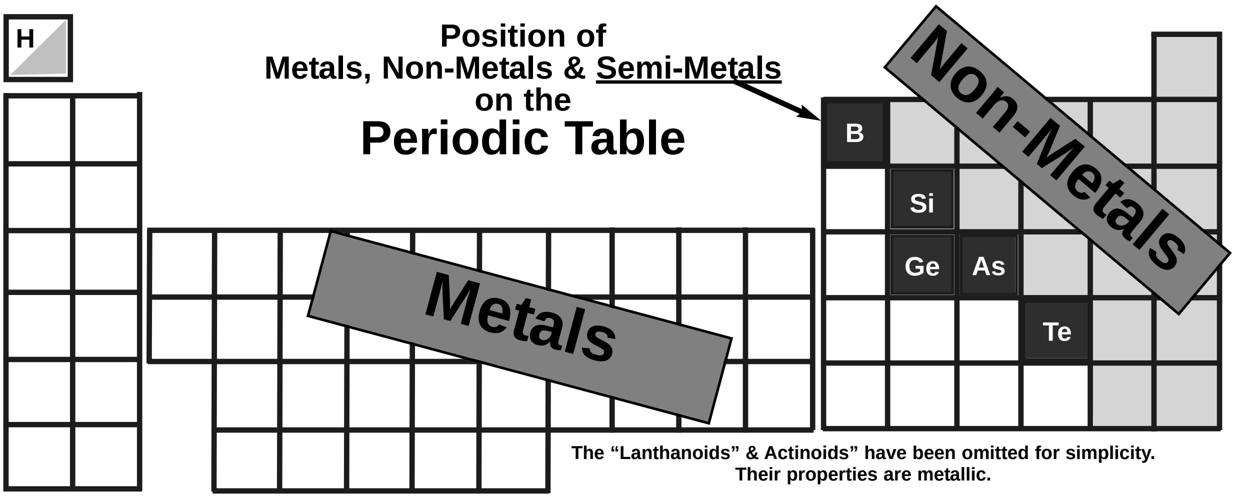
Each element has its own Atomic Number, and its own unique set of properties. However, most elements fall into just 2 general categories...

METALS		NON-METALS
Shiny	<appearance>	Dull (most)
Most hard solids	<state at 25^{o}C>	Some solids, many gases
High	<conductivity> (of electricity)	Low (except carbon)
Malleable	<can be rolled into sheets>	Brittle... shatter
Ductile	<can be pulled into wires>	Brittle... snap

The Semi-Metals (“Metalloids”)

There is also a small group of elements which have properties that are “in-between” and do not fit clearly into the metal or non-metal classification. This group includes Silicon and Germanium which have properties as follows:

- **Shiny appearance, but glass-like rather than metallic.**
- **Hard solids at 25^{o}C (i.e. high melting point).**
- **Semi-conductors of electricity (in-between).**
- **Brittle (non-malleable, non-ductile).**



Position of Metals, Non-Metals & Semi-Metals on the Periodic Table

The “Lanthanoids” & Actinoids” have been omitted for simplicity.
Their properties are metallic.

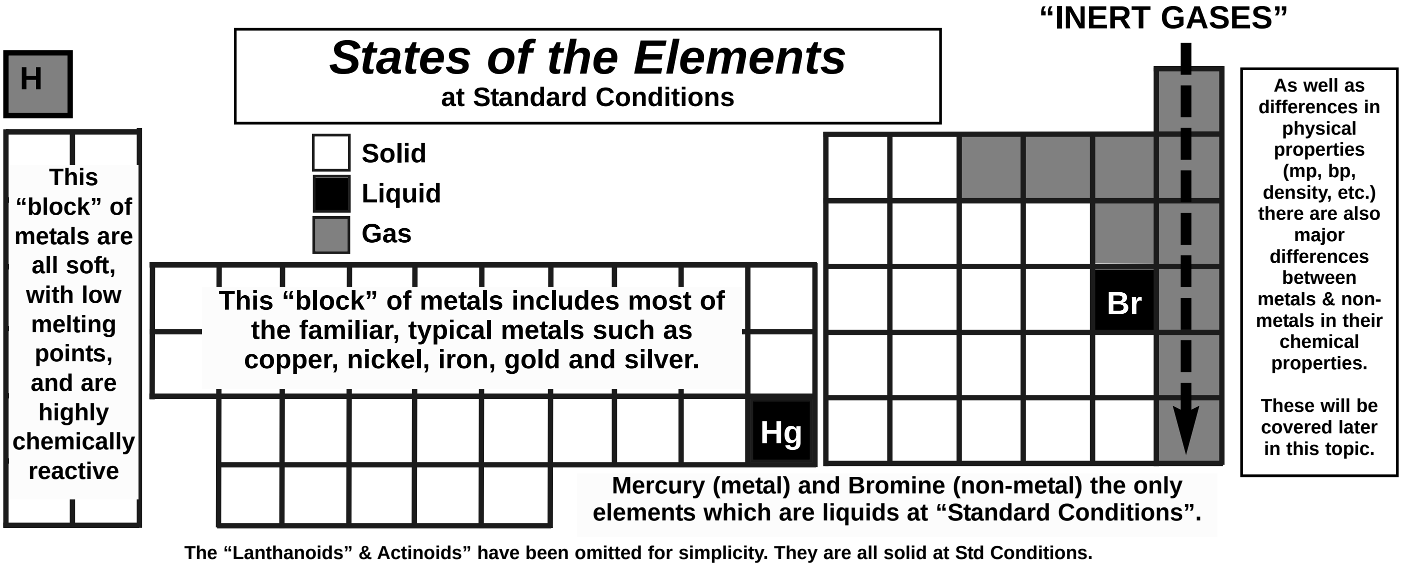
The States of the Elements

You need to understand that whether a substance is solid, liquid or gas is determined by its melting point (m.p.) and boiling point (b.p.).

For example, consider these:

Element	m.p.(^{o}C)	b.p.(^{o}C)	State at 25^{o}C
Iron (Fe)	1535	3000	solid
Mercury (Hg)	-39	357	liquid
Oxygen (O_2)	-219	-183	gas

Changing the pressure changes the mp. and b.p., so that’s why we specify a pressure as well as a temperature when describing what state a substance is. In fact, 25^{o}C and a nominated pressure close to the average atmospheric pressure is known as “Standard Conditions” and is the set of conditions under which chemical measurements are usually made and properties described.



States of the Elements
at Standard Conditions

☐ Solid
■ Liquid
▨ Gas

This “block” of metals are all soft, with low melting points, and are highly chemically reactive

This “block” of metals includes most of the familiar, typical metals such as copper, nickel, iron, gold and silver.

Mercury (metal) and Bromine (non-metal) the only elements which are liquids at “Standard Conditions”.

As well as differences in physical properties (mp, bp, density, etc.) there are also major differences between metals & non-metals in their chemical properties.

These will be covered later in this topic.

The “Lanthanoids” & Actinoids” have been omitted for simplicity. They are all solid at Std Conditions.

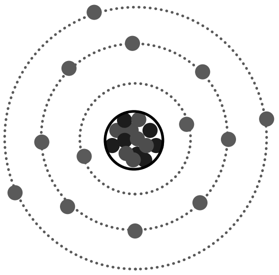
Electron Orbits

The following pages cover ideas you may have studied previously. Use this as revision.

The Concept of Orbits

When the structure of atoms was first becoming understood, (in the early 20th century) it was firstly imagined that the electrons whizzed around the nucleus like a swarm of demented bees. There was no organisation or pattern to their orbital paths.

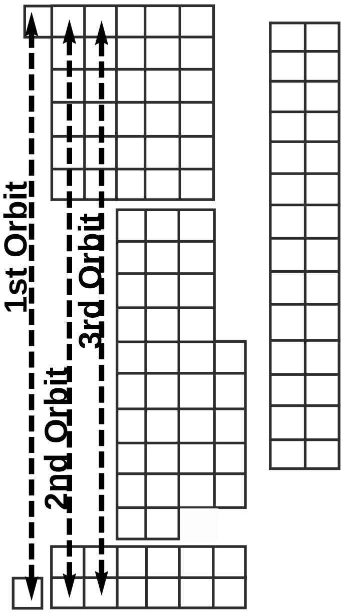
This idea didn't last very long, for reasons which will be covered a little later.



Orbital Layers

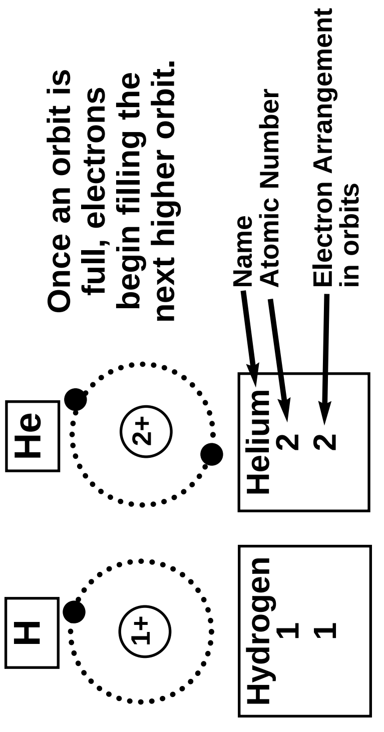
The original idea was soon modified to include the idea that the electrons were arranged in different “levels” or “shells” at different distances from the nucleus.

Each orbit can only hold a certain maximum number of electrons. Chemistry is really all about the exact arrangement of electrons in their orbits, in particular the outer-most orbit of that atom.



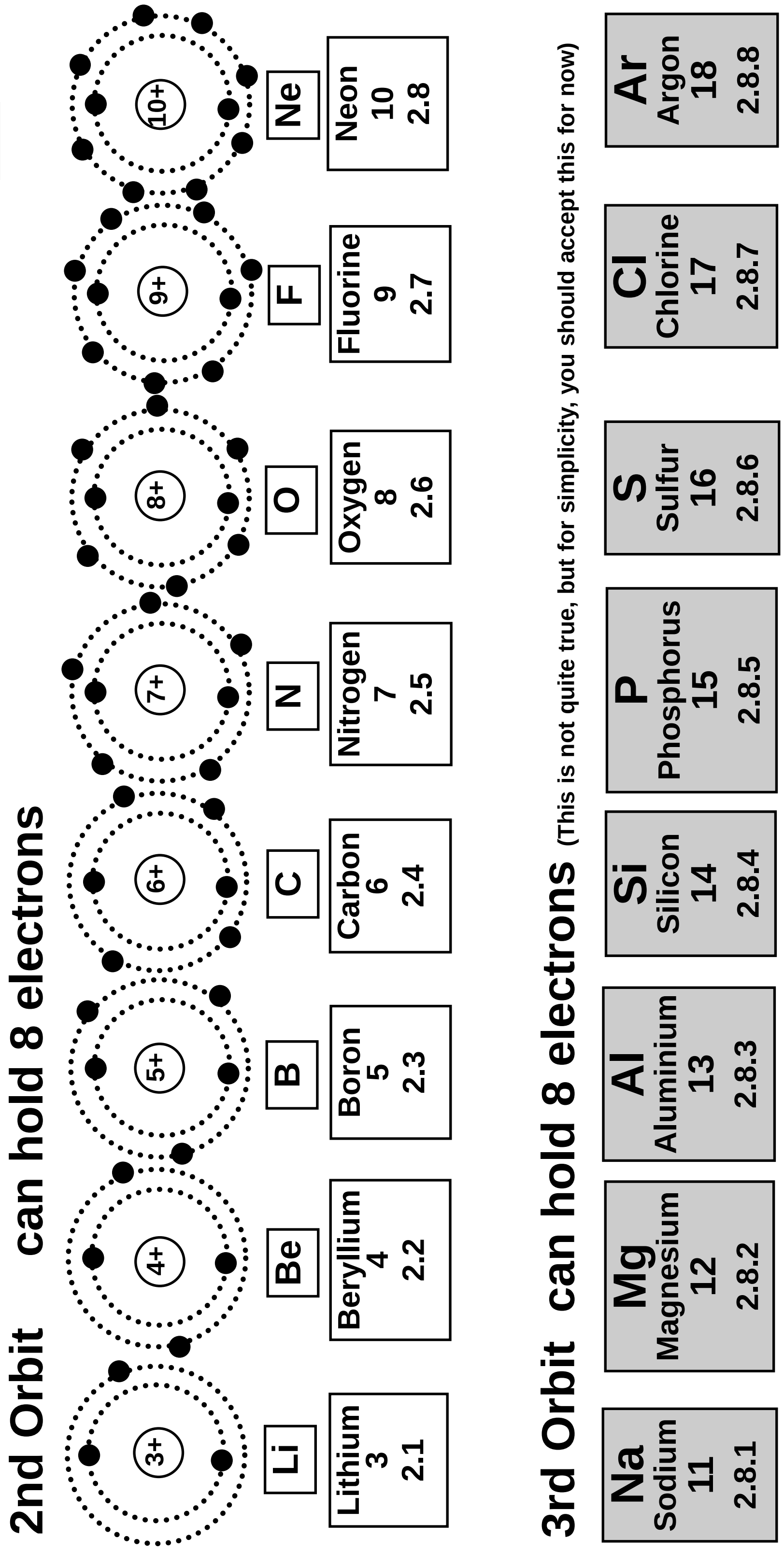
1st Orbit can hold only 2 electrons

Once an orbit is full, electrons begin filling the next higher orbit.



2nd Orbit can hold 8 electrons

3rd Orbit can hold 8 electrons (This is not quite true, but for simplicity, you should accept this for now)

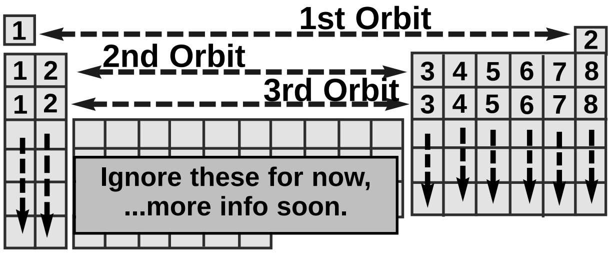


Electrons in the Outer Orbit

If you consider just the number of electrons in the outer orbit of each element, another pattern appears on the Periodic Table.
At right are the data for the first 18 elements.

See the pattern?
Each period corresponds to one orbit of electrons AND elements in the same column have the same number of electrons in their outer orbit.

No. of Electrons in Outer Orbit



For the elements in the far right column (Inert Gases) you should note that their outer orbit is full.

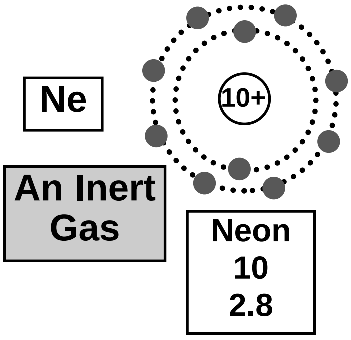
Each row of the Periodic Table lists elements which are filling the same orbit.
Each column lists elements which have the same number of electrons in their outer orbit.

The Importance of a Full Outer Orbit

Quantum Energy Levels

The orbits of the electrons around the nucleus are not just places for electrons to hang out, they are "Quantum Energy Levels" within the atom.

In the strange world of quantum energy, an atom achieves great energy stability if its outer orbit is full.



All the Inert Gas elements already have a full outer orbit.
They are very stable and have no need to do anything to become "perfect".

That is why they do not react with anything and do not form compounds.

How Atoms Get a Full Outer Orbit

All the other elements do NOT have a full outer orbit. To achieve the "perfect" stable energy state atoms can either:

- **GAIN ELECTRONS to fill up their outer orbit.**

OR

- **LOSE ELECTRONS and completely shed their outer orbit.**

Their "new outer orbit" becomes the one underneath, which is full and "perfect".

As you will see, to gain or lose electrons, atoms must interact with each other.

The result is Chemistry!

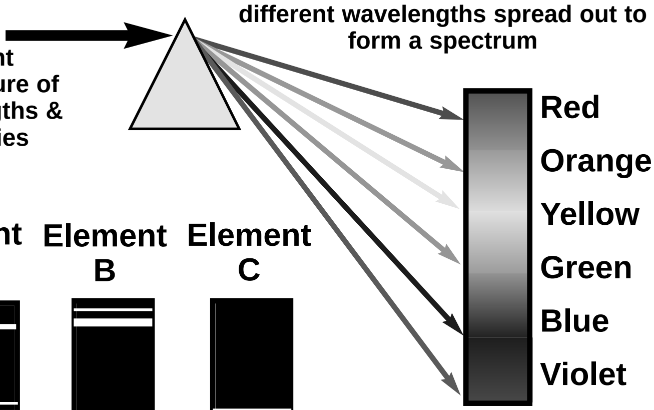
The Bohr Model of the Atom

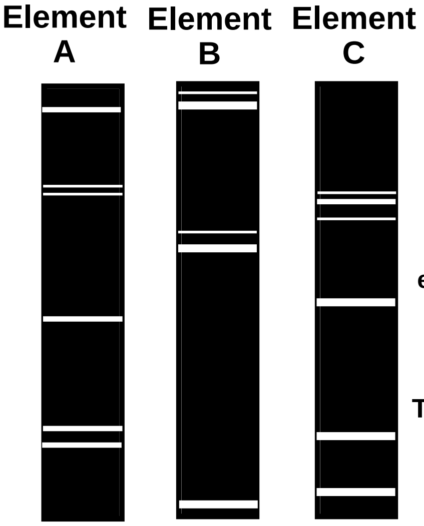
Niels Bohr
1885-1962

The model of the atom used in the previous slides is known as the "Bohr model" because it was proposed by the Danish scientist Niels Bohr in 1913. His model was eagerly accepted at the time because it overcame a theoretical problem with the previous idea of atomic structure AND it explained something that Science had been grappling with for decades prior...

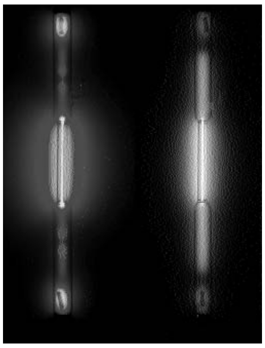
...Emission Spectra

You should be familiar with the idea of a "spectrum" of light. For example, if "white" light is passed through a prism, the different wavelengths are separated, and the familiar rainbow colours appear.





<u>Spectral lines</u>
Each line is light of one exact wavelength. Light is only emitted at certain precise wavelengths



In the 19th century, it was discovered that a tube of any gas would glow when high-voltage electricity was applied.

Such tubes are called "discharge tubes".

If a discharge tube is filled with a <u>pure element</u> (gas), the spectrum shows very narrow bright lines on a dark background. (diagrams at left)

This is because only certain wavelengths are given out. The pattern of lines is characteristic for each element & can actually be used to identify elements. (That's how we know what the stars are made from.)

The lines in the emission spectrum of <u>hydrogen</u> had been discovered some 30 years before Bohr's theory. Each line was given a name (α, β, χ and δ) & the precise wavelength of each had been measured.

No-one could explain them, but mathematicians had worked out that the exact wavelengths of the hydrogen spectrum lines could be calculated from a made-up equation (called the "Rhydberg Equation") which contained a series of consecutive whole numbers.

The fact that the equation worked perfectly was strong evidence that there was an underlying "law" controlling the spectral lines. The fact that a series of integer numbers were involved was a clue that connected the whole thing to Plank's recently invented Quantum Theory.

Investigating the Evidence: possible prac.work

Flame Tests

You may carry out some prac. work to investigate spectral lines.

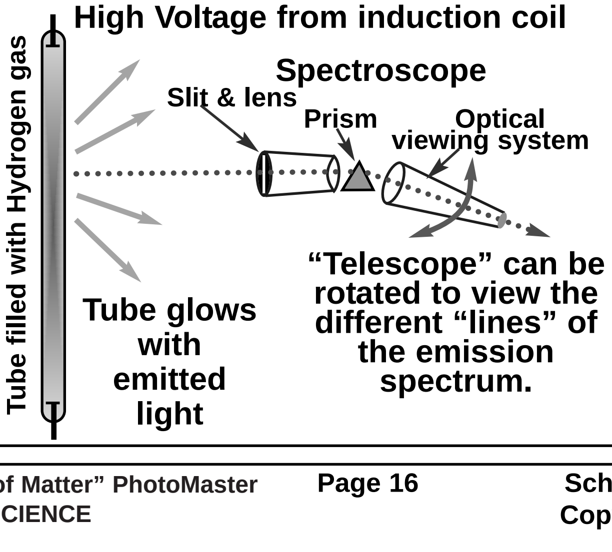
The simplest experiments involve "<u>Flame Tests</u>" in which samples containing certain metals "flare" with characteristic colours when energised in a bunsen flame.

The colours are due to precise wavelengths of light emitted by energised electrons as they drop back to lower energy levels within the atoms.

Colours can be used to identify some metals.

Emission Spectrum of Hydrogen

You may have observed the emission spectrum for hydrogen by using a <u>spectroscope</u> to view the light from a <u>discharge tube</u> filled with low-pressure hydrogen gas.



You will have seen that the light from a hydrogen discharge tube is composed of 4 visible bright lines of light.

Each line is one single wavelength of light.

The Bohr Model of the Atom (cont.)

Quantum Theory

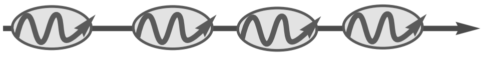
In 1900, Max Plank had proposed the Quantum Theory to explain the details of "Black-Body Radiation Curves". To explain things, he proposed that energy, such as light, came in whole-number multiples of fixed "units" called quanta. (singular = "quantum")

An analogy is to think about our cash money. The smallest possible amount of cash you can have is a 5c coin. This is the "quantum of cash". No matter how much cash you have, it must be a multiple of the quantum amount. You can't have 37c in cash!

Plank was saying that light cannot have just any quantity of energy (like 37c); it must be a multiple of the basic quantum.

Plank used this idea purely as a clever "mathematical trick" to help explain something that could not be explained any other way, but other scientists found this idea useful to solve other difficult problems.

In 1905, Einstein explained the strange phenomenon of the Photoelectric Effect by using Plank's quantum idea. To do this, he proposed that light is not just a wave, nor a stream of particles, but made up of "wave packets" which have BOTH wave characteristics AND particle-like properties.

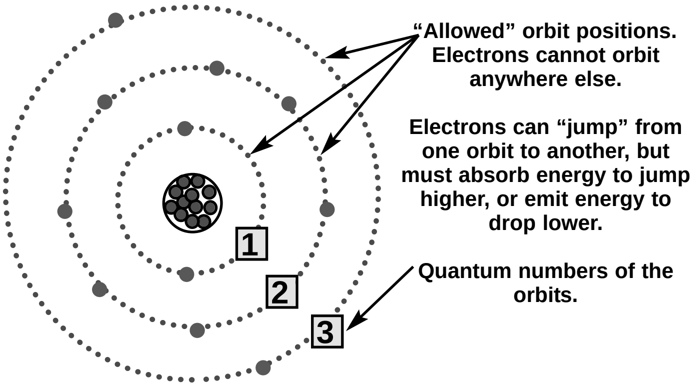


Light is a stream of "wave packets"... "PHOTONS".
They have wave properties such as refraction & interference.
They can also behave like a particle sometimes. Each photon carries an amount of energy which is an exact multiple of Plank's basic quantum of energy.
Each photon is both a particle AND a wave!

Bohr's Suggestions

"Allowed" Orbits for Electrons

The previous atomic model had imagined that the electron orbits were more or less random. Bohr theorised that there are a series of orbits, at fixed distances from the nucleus and each orbit can hold certain maximum number of electrons.



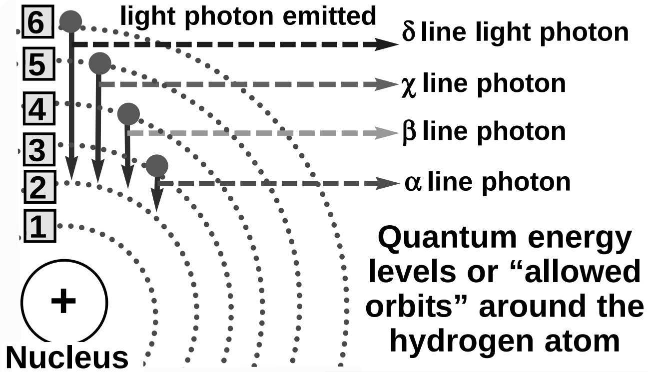
Electrons Gain or Lose Energy to "Jump"

To jump up to a higher orbit, an electron must gain a certain quantity of energy. If it drops back to a lower orbit, it must emit that exact same amount of energy. These quantities of energy are "quantised", so each orbit is really a "quantum energy level" within the atom.

The amount of energy absorbed or emitted during a "jump" is defined by Plank's Quantum Equation, and the corresponding wavelengths of light are defined by the Rhydberg Equation. The integer numbers in the equation turn out to be the "quantum numbers" of the orbits, counting outwards from the nucleus.

Bohr's Theory & the Hydrogen Emission Lines

Bohr's atomic model explained the mystery of the spectral lines, not just in general terms, but in precise mathematical detail. That's why the model was accepted and is still the model used in schools today.



Bohr showed that the Hydrogen α emission line was due to an electron dropping from the 3rd orbit down to the 2nd orbit. It must lose a precise amount of energy, so it emits a photon of light at an exact wavelength which can be calculated. This calculated wavelength agreed perfectly with the observed spectral line. Plank's Quantum Equation can calculate the energy of that photon of light. Bohr argued that this amount of energy must represent the difference in energy from orbit 2 to orbit 3.

The other hydrogen spectral lines must be due to electrons dropping from higher orbits into the 2nd orbit and all the calculations work out perfectly!

It all worked! Bohr's idea gave a theoretical explanation for the Rhydberg Equation, which had been empirically derived to describe the spectral lines.

Schrodinger & Orbital Theory

Quantum Mechanics

Bohr's model came out in 1913. Then, because of WW1, there was a pause in scientific development. However, from the mid-1920's the "Quantum Theory" was fleshed-out to form "Quantum Mechanics", a comprehensive Science to fully explain what is happening at the atomic level.

There were many great scientists involved. If interested, you should research the names DeBroglie, Pauli & Heisenberg (and these may lead you to many more).

Erwin Schrodinger in 1933

However, the scientist who became known as the "father of Quantum Mechanics" was the Austrian, Erwin Schrodinger.

Schrodinger's mathematical "wave function" equation became the basis for understanding many aspects of the sub-atomic world. The Science of "Quantum Mechanics" now under-pins much of modern Physics & Chemistry.

In the brief summary of ideas which follows, you should bear in mind a statement attributed to another great scientist, Richard Feynman, who said:

"If you think you understand Quantum Mechanics, then you don't understand Quantum Mechanics!"

(In fact, it may be an urban myth that he ever said this... but, if not, he should have!)

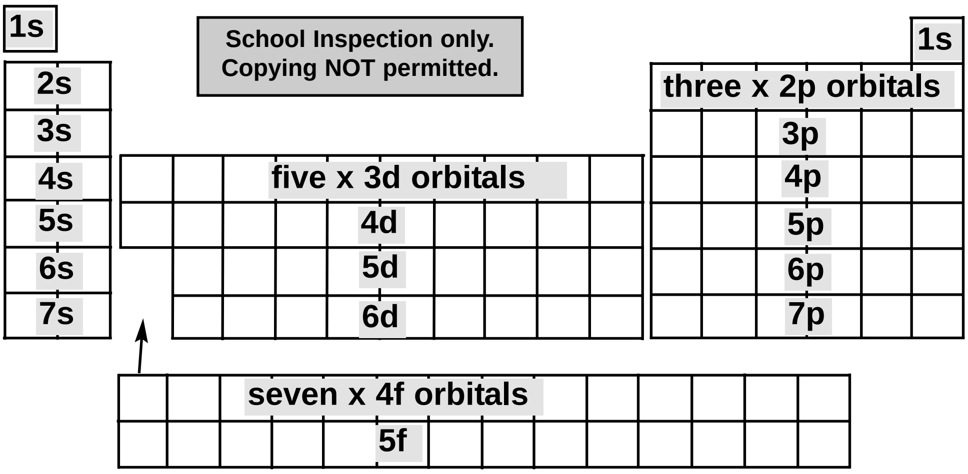
Atomic Theory & Quantum Mechanics

The Bohr model has not been totally replaced by developments in Quantum Mechanics, but has certainly been modified & added to. The full story of the electron orbits can now be summarised as follows:

- **Electrons in atoms are not particle-like. They act as wave-packets (similar to Einstein's idea about photons of light in 1905). Electrons have a wavelength like a wave & must be thought of as "particle-waves".**

- **Each of Bohr's orbits really contains a number of sub-orbits or "orbitals". Each orbital can contain a maximum of 2 electrons. The orbitals exist because the energy of an electron is quantised in at least 4 different ways. This results in 4 different types of orbitals called s, p, d & f. (Familiar?)**

The relationship between the orbital patterns & the Periodic Table is shown below.



**School Inspection only.
Copying NOT permitted.**

More About the Orbitals

s, p, d & f-blocks

**were named earlier.
Now you can understand the size of each block.**

There is only one "s-type" orbital in each main orbit. It holds just 2 electrons, so the "s-block" is 2 columns wide.

There are three "p-type" orbitals (2 electrons each) so the "p-block" is 6 columns wide.

Five d-orbitals = 10 wide.

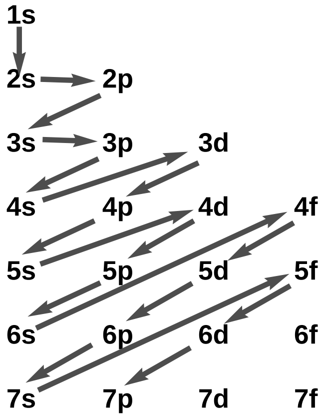
Seven f-orbitals = 14 wide.

Order of Filling the Orbitals

Because some orbitals in adjacent main orbits "overlap" in energy terms, the order of filling-up is not as you might expect.

The diagram shows the order in which orbitals are filled as you move through the Periodic Table.

Note that the modern table of 118 elements fills as far as the 7p orbital only.



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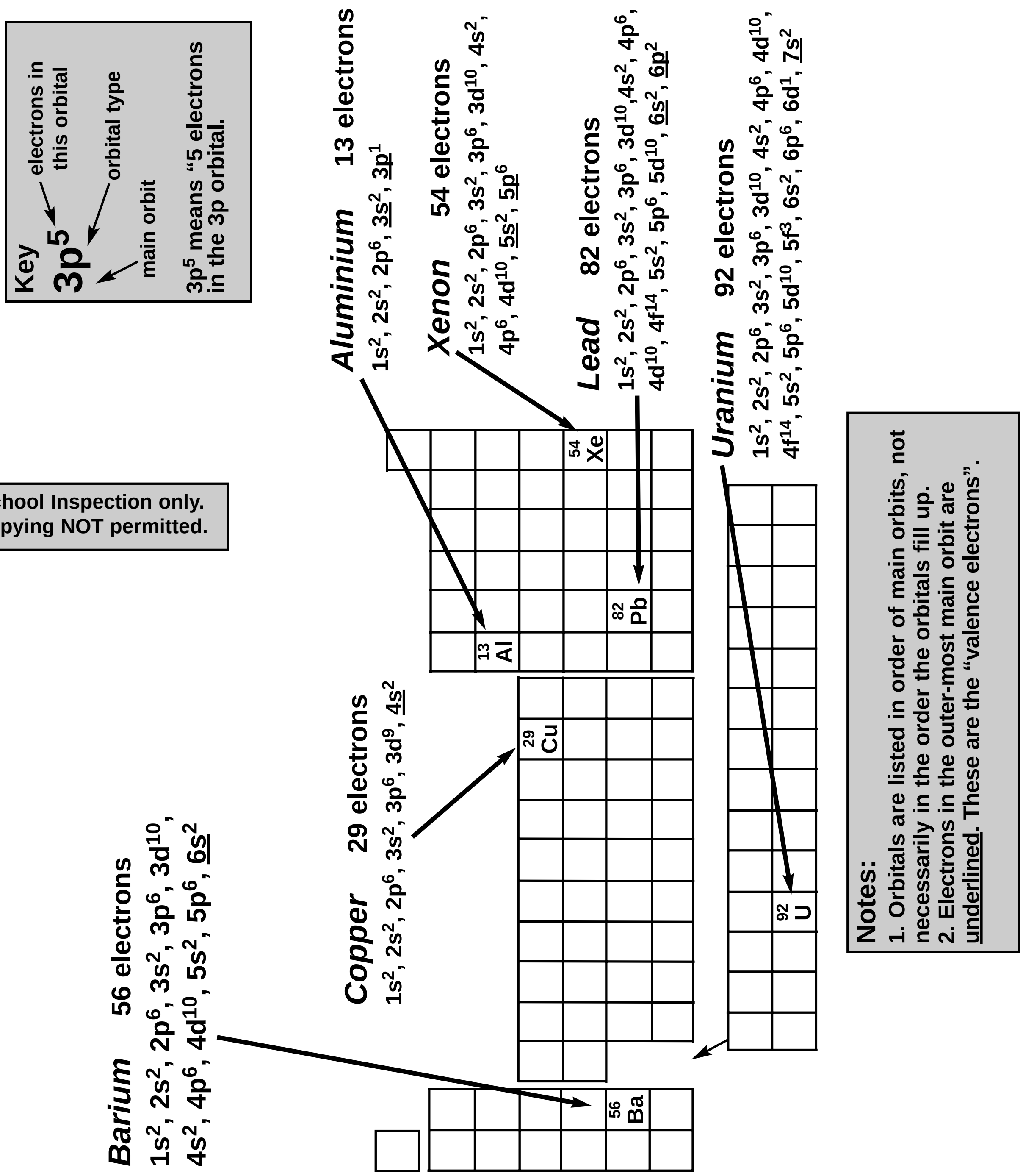
Orbital Notation of an Element

Here, by example, you can learn how each element can be described in “Orbital Notation”. Study each example to get the idea. Study in order of increasing Atomic Number. (start with aluminium)



Key

$3p^5$ — main orbit; orbital type; electrons in this orbital

$3p^5$ means “5 electrons in the 3p orbital.

Aluminium **13 electrons**

$1s^2, 2s^2\ 2p^6, \underline{3s}^2, \underline{3p}^1$

Xenon **54 electrons**

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, \underline{5s}^2, \underline{5p}^6$

Lead **82 electrons**

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 4f^{14}, 5s^2, 5p^6, 5d^{10}, \underline{6s}^2, \underline{6p}^2$

Uranium **92 electrons**

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 4f^{14}, 5s^2, 5p^6, 5d^{10}, 5f^3, 6s^2, 6p^6, 6d^1, \underline{7s}^2$

Copper **29 electrons**

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^9, \underline{4s}^2$

Barium **56 electrons**

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 5s^2, 5p^6, \underline{6s}^2$

Notes:

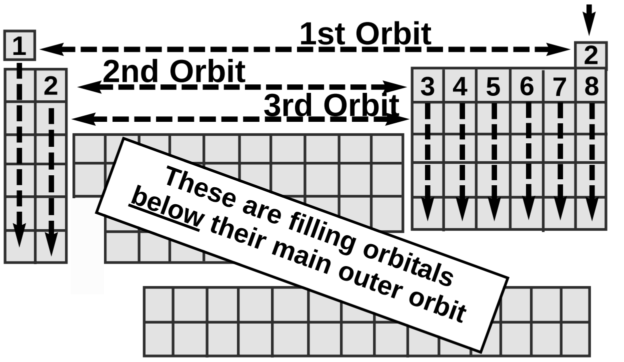
1. Orbitals are listed in order of main orbits, not necessarily in the order the orbitals fill up.
2. Electrons in the outer-most main orbit are underlined. These are the “valence electrons”.

And now, guess what? To understand most of HSC Chemistry, you don’t need to know about this! From here on the Bohr model is enough to know about. Except for a few “tricky bits” involving elements in the d-block, we might rarely mention orbitals again. BUT, it’s in the syllabus, so you need to be aware of it.

Now we get back to *Patterns in the Periodic Table*

Electrons in the Outer Orbit

You may already be aware, from previous studies, of the importance of the number of electrons in an atom's outer main orbit. Outer orbit electrons pretty much control chemical behaviour. Back on page 15, you saw this diagram showing a pattern of the Periodic Table.



Now that you know a little about orbitals, you will see that the outer-most orbitals are ALWAYS either s-type or p-type orbitals. Elements in the d-block & f-block are always filling orbitals in a main orbit below the outer-most. Their outermost electrons are in the s-type orbital at the beginning of that period.

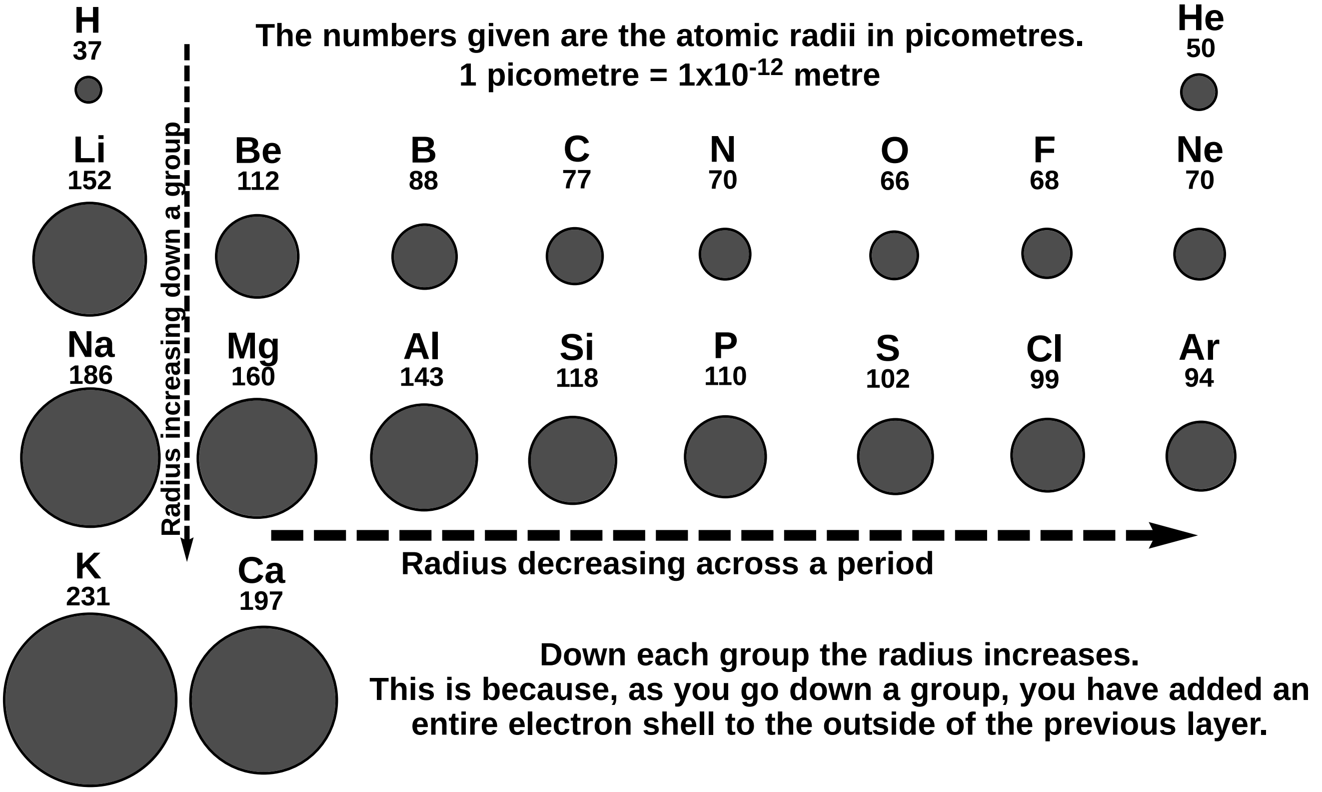
Generally, you may assume that all d-block & f-block elements have 2 electrons in their outer orbit. However, some of the orbitals are so close to each other (in quantum energy terms) that these atoms shuffle electrons between orbitals and become rather unpredictable.

More about that later. For now, learn the pattern above. Important!

Atomic Radius

The size of an atom is the distance across its outer electron shell. You might think that the atoms along each period would be the same size, because it's the same electron orbit being added to. However, the increasing amount of positive charge in the nucleus pulls that orbit inwards closer and closer to the centre.

The diagrams are to scale and show the relative sizes of the first 20 elements



The numbers given are the atomic radii in picometres.
1 picometre = 1×10^{-12} metre

Down each group the radius increases.
This is because, as you go down a group, you have added an entire electron shell to the outside of the previous layer.

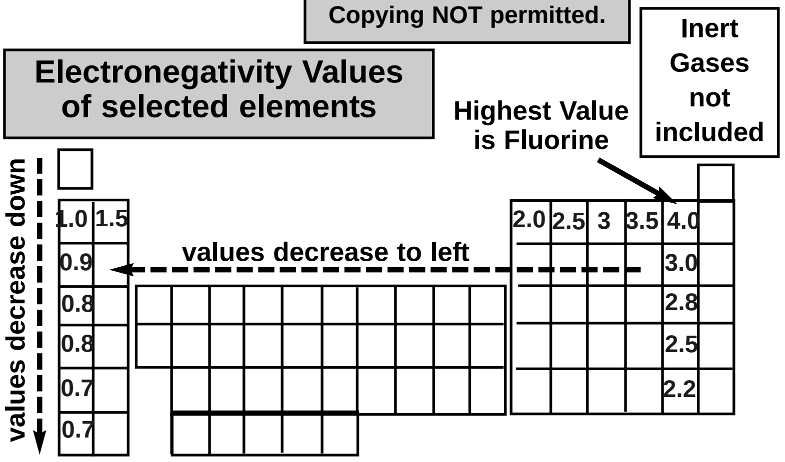
More Patterns in the Periodic Table

Electronegativity

is a value assigned to each element to describe the power of an atom to attract electrons to itself. Electronegativity will become very important later when you study chemical bonding.

Atoms with a tendency to gain electrons and form negative ions have high values. Atoms with a tendency to lose electrons easily and form positive ions have very low values.
(If you haven't yet learnt about ions, skip this & come back later.)

Electronegativity Values of selected elements



Once again, there is a pattern in these values in the Periodic Table.

Ionisation Energy

The Ionisation Energy of an element is the energy required to remove an electron from an atom.

For technical reasons, the measurement of this energy is carried out for atoms in the gas state.

$$A_{(g)} \longrightarrow A^{+}_{(g)} + e^{-}$$

where "A" stands for any atom in the gas state

We know that zinc atoms normally lose 2 electrons to form the Zn^{2+} ion. However, the formal definition for this process involves just the loss of 1 electron.

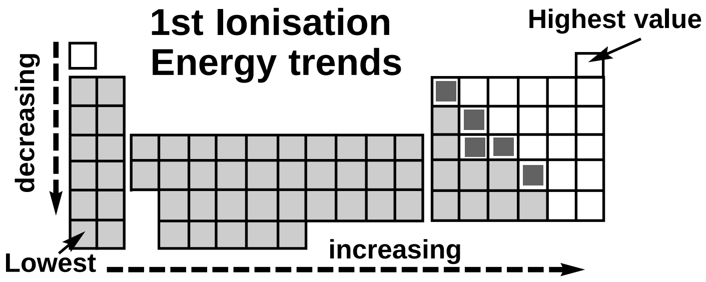
$$Zn_{(g)} \longrightarrow Zn^{+}_{(g)} + e^{-}$$

The energy required for this to happen is the "1st Ionisation Energy".

Every element has its own characteristic value, even those elements which would not normally lose electrons, such as non-metals like chlorine.

$$Cl_{(g)} \longrightarrow Cl^{+}_{(g)} + e^{-}$$

**Normally a chlorine atom forms a negative ion by gaining an electron. Technically though, it is possible for it to lose an electron if enough energy is added.
This energy is the "1st Ionisation Energy".**



Explanations

**1st I.E. increases to the right because each atom across a period has more and more (+ve) nuclear charge attracting and holding electrons in the orbit concerned.
Therefore, it requires more energy to remove an electron.**

**1st I.E. decreases down each group because, at each step down, an extra whole layer of electrons has been added to the outside of the atom. The outer shell is further away from the nucleus, and is partially "shielded" from nuclear attraction by the layers of electrons underneath it.
Therefore, it becomes easier and easier to remove an electron.**

One Last Pattern... Chemical Reactivity

What is meant by "reactivity" is difficult to define. It could be defined by the "ease" of starting a reaction (linked to a measurement called "activation energy") or by the rate at which the reaction proceeds, or by the violence & energy release. Futhermore, reactivity very much depends on which other substance is involved. Here we will summarise some general patterns covering many of these factors.

Group 16: Oxygen's Group

Oxygen is highly reactive. It will react with nearly every other element except the inert gases. Oxygen is involved in any reaction involving "burning" and in the corrosion of metals. Its importance to the chemistry of living things is well known.

An entire category of chemical reactions is called "oxidation" because of oxygen's wide-ranging chemical activity.

Reactivity decreases down this group.

Halogens (Group 17)

These elements are the most reactive of the non-metals. They have high values of electronegativity, which means they readily "steal" an electron from other chemicals to fill their almost-full outer orbit. They react rapidly with metals & with biological chemicals. This makes them useful as bleaches & disinfectants. It also makes them very dangerous!

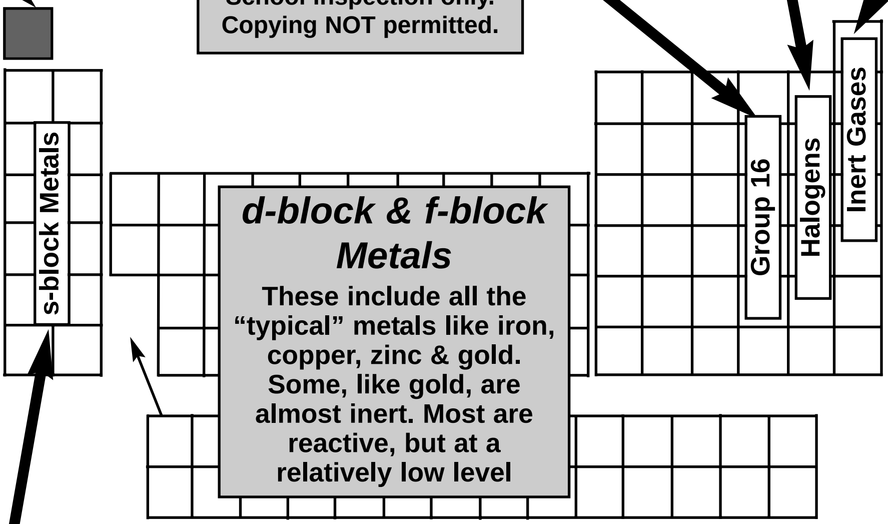
Reactivity decreases down the group.

Hydrogen

As always, Hydrogen is in a category all alone. It can react explosively with oxygen, but the reaction is not easy to start. It can form comounds with a variety of metals & non-metals, but is notable for its myriad of compounds with carbon.

Inert Gases

All have full outer electron orbits, so they DO NOT react chemically. Reactivity = zero.



d-block & f-block Metals

These include all the "typical" metals like iron, copper, zinc & gold. Some, like gold, are almost inert. Most are reactive, but at a relatively low level

s-block Metals Groups 1 & 2

These are the most reactive of the metals. They will react readily (sometimes violently) with oxygen, water or acids. They will react with dissolved compounds of other metals in order to "displace" the other metal from the compound. Reactivity increases down the group & decreases to the right.

General Pattern in the Non-Metals

Chemical activity:

- **decreases down each group.**
- **decreases to the left.**

General Pattern in the Metals

Chemical activity:

- **decreases to the right.**
- **increases down each group.**
(but exceptions occur in the d-block)

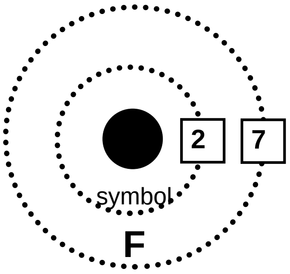
3. Chemical Bonding

Chemical reactions occur because every atom achieves its best possible energy state (from Quantum Mechanics) if it has a full outer orbit of electrons. In effect, this means 8 electrons (2 in s-type & 6 in p-type orbitals) in the outer-most orbit (2 for elements of the 1st Period). The only elements which already have this “magic number” are the Inert Gases of Group 18. All other elements can achieve “perfection” by exchanging or sharing electrons. This occurs during a chemical reaction and creates a chemical bond.
This is how elements combine into compounds.

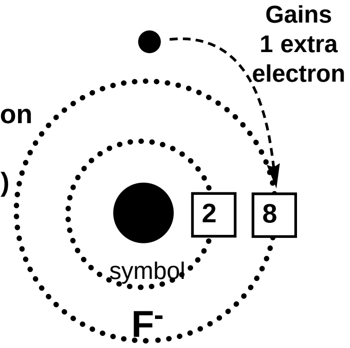
Formation of Ions

Most atoms do not have the correct number of electrons to fill their outer “shell” or orbit. However, many atoms will readily lose or gain electrons in order to achieve this. An atom which has gained or lost electron(s) is called an ION.

Example 1: Formation of a Fluoride Ion



An atom of Fluorine has
9 protons (+)
10 neutrons
9 electrons (-)
Electron Configuration = 2.7



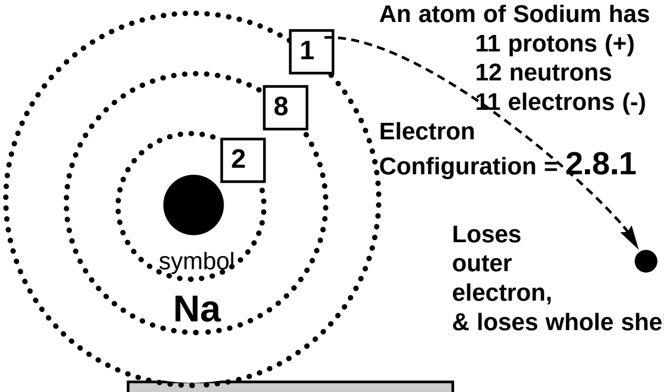
If this atom gains an electron it becomes a Fluoride Ion (F^-). (note change in name)

Electron Config. = 2.8

It still has 9 protons (+ ve charge) but now has 10 electrons (- ve). Overall, it now has a negative charge.

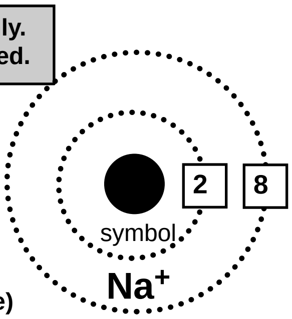
OUTER SHELL IS FULL = BEST ENERGY STATE

Example 2: Formation of a Sodium Ion



An atom of Sodium has
11 protons (+)
12 neutrons
11 electrons (-)
Electron Configuration = 2.8.1

Loses outer electron, & loses whole shell.



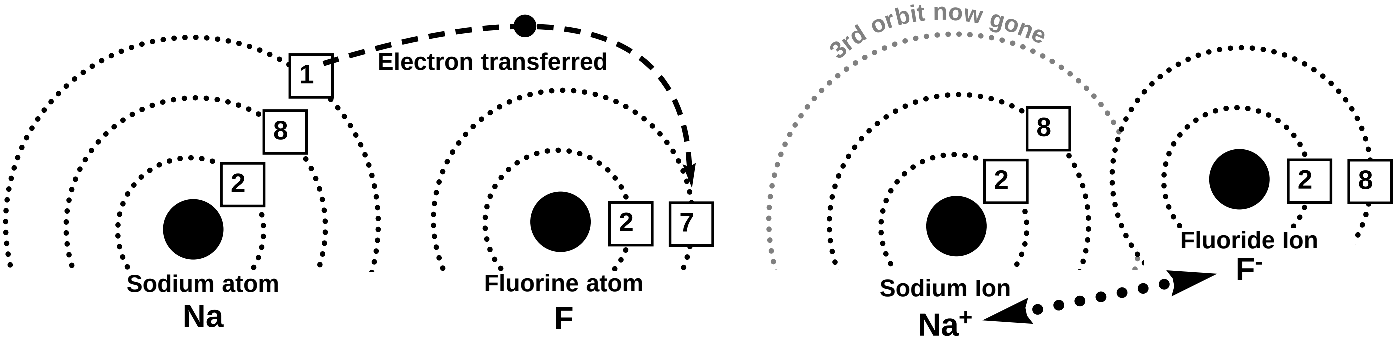
If this atom loses an electron it becomes a Sodium Ion (Na^+). (note NO change in name)
Electron Configuration = 2.8

It still has 11 protons (+ ve charge) but now has 10 electrons (- ve). Overall, it now has a positive charge.
OUTER SHELL IS FULL = BEST ENERGY STATE

(It has lost the 3rd shell entirely, so the full 2nd shell is now its outermost orbit)

Formation of Ionic Bonds & Compounds

If a sodium atom came near to a fluorine atom, it should be obvious (from above) what will happen...



Both atoms become ions...

Opposite electrical charges attract, so these ions must now stick together...

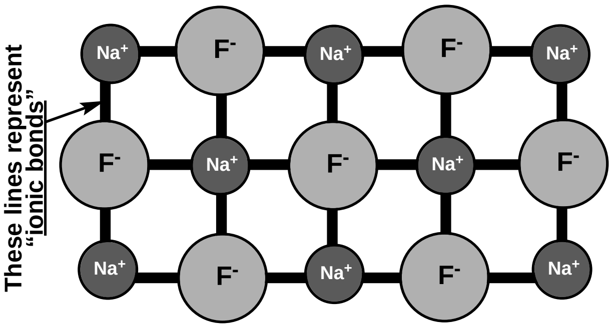
this is the COMPOUND Sodium Fluoride (NaF)

Formation of Ionic Bonds & Compounds *(cont.)*

Ionic Lattices

In fact, of course, you don't just get 1 sodium atom reacting with 1 fluorine atom. In real situations there are billions of atoms. After all the ions have formed, each positive sodium ion is attracted to every nearby fluoride ion, and vice versa.

The result is that you don't just get pairs of opposite ions, but huge, 3-dimensional lattices of +ve and -ve ions.



Each bond is really the attraction between opposite electric charges

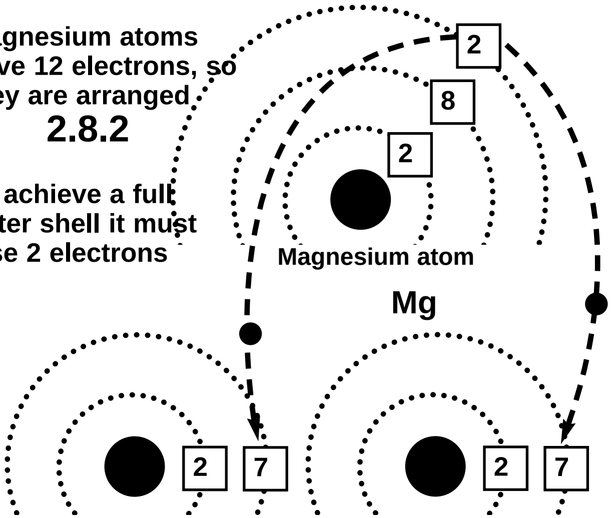
The chemical formula for any ionic compound is an “empirical formula”... it shows only the ratio between the ions, not the actual numbers that are present. In ionic compounds there are no discrete molecules. In the solid state an ionic compound forms a crystal, which is a huge array of billions of ions in a lattice.

Another example of an Ionic Compound...

What if it was magnesium atoms that reacted with fluorine?

Magnesium atoms have 12 electrons, so they are arranged
2.8.2

To achieve a full outer shell it must lose 2 electrons



2 Fluorine atoms
Each one accepts an electron from the magnesium atom

The magnesium loses 2 electrons, so it will now have 2 more (+) protons than electrons. The magnesium ion has a double +ve charge... Mg^{2+}

Two fluoride ions are formed, each one with an extra electron... F^-

The formula for magnesium fluoride is MgF_2 because the ratio between the ions is 1: 2

Ionic Compounds
form when atoms gain, or lose, electrons creating ions.

METALS
always lose electrons forming POSITIVE ions.

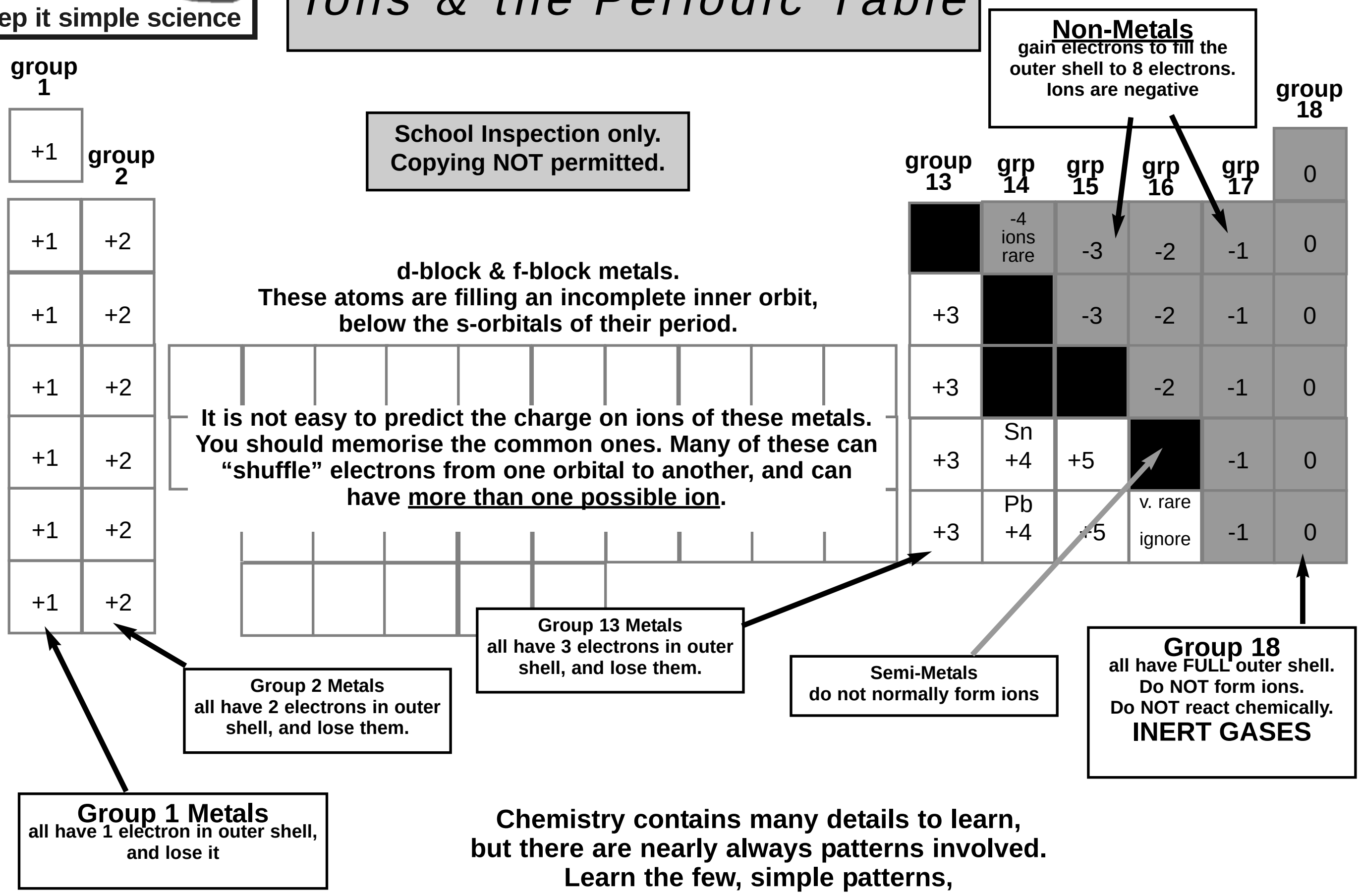
NON-METALS
always gain electrons forming NEGATIVE ions.

The name of the atom changesto end -IDE, for the ion.

Positive & Negative Ions
are attracted electrically, forming ionic crystal lattices.

The chemical formula shows the ratio between ions not the actual numbers involved.
(“Empirical Formula”)

Ions & the Periodic Table



Non-Metals
gain electrons to fill the outer shell to 8 electrons.
Ions are negative

School Inspection only.
Copying NOT permitted.

d-block & f-block metals.
These atoms are filling an incomplete inner orbit, below the s-orbitals of their period.

It is not easy to predict the charge on ions of these metals. You should memorise the common ones. Many of these can "shuffle" electrons from one orbital to another, and can have more than one possible ion.

Group 13 Metals
all have 3 electrons in outer shell, and lose them.

Group 2 Metals
all have 2 electrons in outer shell, and lose them.

Semi-Metals
do not normally form ions

Group 18
all have FULL outer shell.
Do NOT form ions.
Do NOT react chemically.
INERT GASES

Group 1 Metals
all have 1 electron in outer shell, and lose it

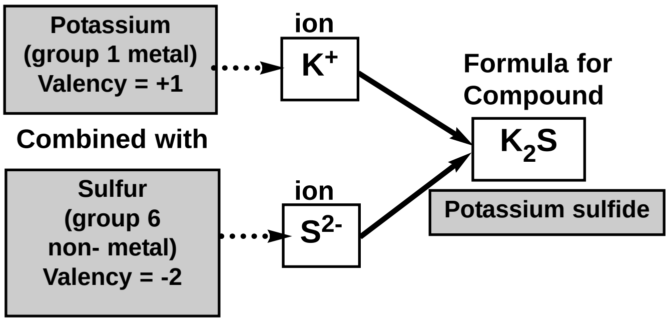
Chemistry contains many details to learn, but there are nearly always patterns involved. Learn the few, simple patterns, rather than memorising masses of detail.

Valency

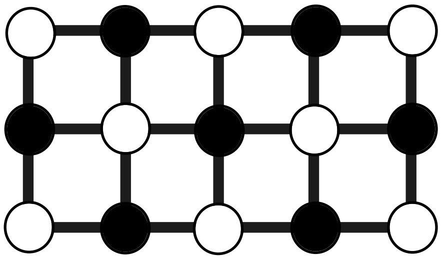
The "valency" of a chemical species refers to its "combining ratio" with other species.

For the simple ions, the charge on the ion (including sign) is the same as valency.

If you know (or can figure out by the patterns) the valency of species, you can predict the chemical formula of the compound. Study these examples:

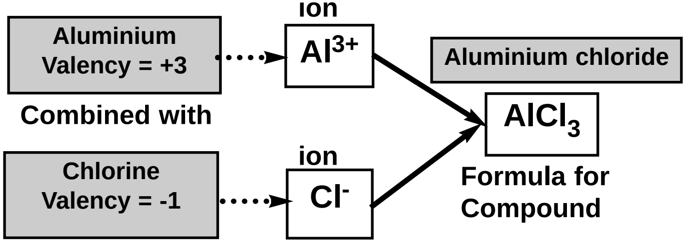


You must have same amount of (+) and (-) electric charge, in every case.



Model of an Ionic Crystal Lattice.

An ionic formula is "Empirical" and describes the ratio of ions.



More Than Just Simple Ions

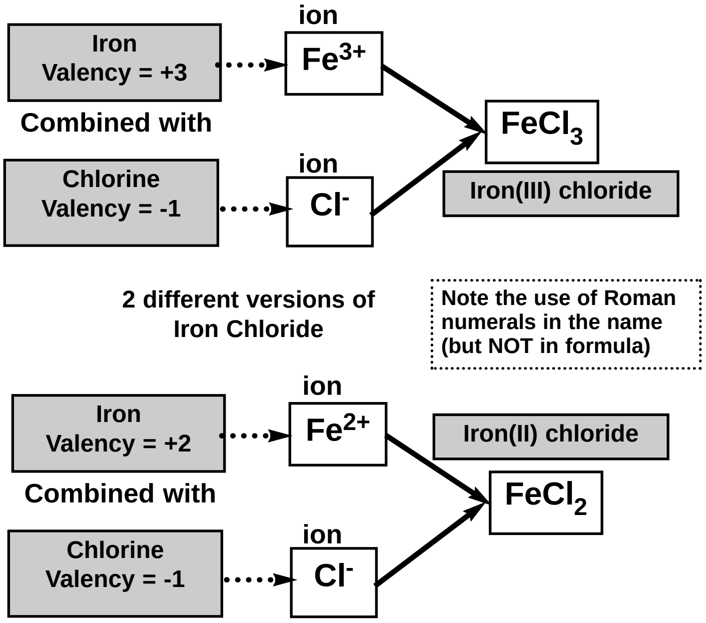
As much as we'd like to keep it really simple, there are a few complications with ions that you must know about.

Multi-Valency Metals

Many of the metals belonging to the d-block (& f-block) of the Periodic Table can form ions in more than one way.

For example, atoms of Iron (Fe) most commonly have 3 electrons in the outermost electron shell. To form ions, the atoms lose these 3 electrons and thereby form Fe^{3+} ions.

However, sometimes the iron atom can "shuffle" its electrons between its outer shell (the 4th energy level) and the incomplete 3rd shell in such a way that it has only 2 electrons in the outer shell. In this situation the atoms will lose only 2 electrons to form an ion... Fe^{2+} ions form.



Note the use of Roman numerals in the name (but NOT in formula)

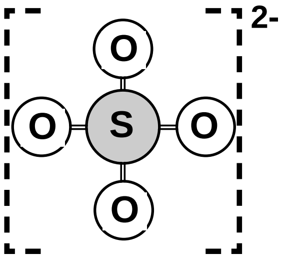
Polyatomic Ions

As well as the simple ions which form when individual atoms gain or lose electrons, there are a number of more complicated ionic species you must know about because they are very common, and cannot be avoided.

These are the "polyatomic" ions (poly=many) which are composed of a group of atoms which have an ionic charge on the whole group, due to the gain or loss of electron(s). The entire group acts chemically just like a single, simple ion, and can join with other ions forming compounds and ionic crystal lattices.

One common example is the "sulfate" ion,

SO_4^{2-}

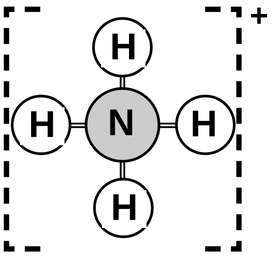


Somewhere within this group of 5 atoms there are <u>2 extra electrons</u>, in excess of the total protons these atoms contain. The ion has a valency of -2.

Other examples are:

- **Nitrate (NO_3^-) ion (valency -1)**
- **Hydroxide (OH^-) ion (valency -1)**
- **Carbonate (CO_3^{2-}) ion (valency -2)**

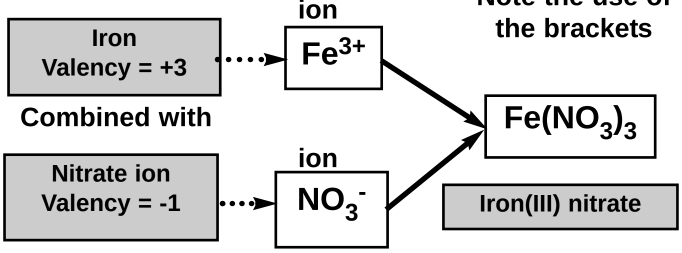
Most of the common polyatomic ions have (-ve) charge and valency. Only one common example has a (+ve) valency like a metal. This is the ammonium ion (NH_4^+).



This group consists of a nitrogen atom and 4 hydrogens. Compared to the total protons, this group has <u>one less electron</u> so it acts as an ion with a valency of +1.

Working out a chemical formula is done exactly as before, except names do NOT change and brackets are needed when 2 or more polyatomic groups are involved.

<u>Example</u>:



Note the use of the brackets

A Summary:
Formulas & Names for Ionic Compounds

Formulas

1. Determine the (+ve) and (-ve) ion involved.

2. Work out the <u>minimum</u> number of each ion which gives <u>equal amounts</u> of (+ve) & (-ve) charge.
   Example: you need 3x(-1) to match (+3)

3. Write symbol for the (+ve) ion first.

4. Use sub-scripts to show ratio of ions. Number “1” is not written.
   e.g. $FeBr_2$

5. If a polyatomic ion is involved;
   - brackets MUST be used if more than one polyatomic ion.
     e.g. $Mg(NO_3)_2$
   - bracket must NOT be used if only one polyatomic ion.
     e.g. $NaNO_3$

Note: The symbols for an ion must contain electric charge, written as a super-script. e.g. Fe^{3+}

The formula for a compound must NOT contain electric charges.

A Table Summarising Symbols & Valencies is at the end of these notes

Naming Compounds of Simple Ions

1. Name the (+ve) ion (metal) first. Its name is always the same as element name.

2. Add the name of the (-ve) ion (non-metal), but altered to end in -IDE.
   e.g. oxygen becomes “oxide”
   phosphorus becomes “phosphide”

Naming Compounds of <u>Metals with Multiple Valencies</u>

As above, but (in brackets) write the Roman numeral corresponding to the valency number of the metal ion.

e.g. $FeBr_2$ is “iron(II) bromide” (Fe^{2+} ion)
[speak “iron-2-bromide”]
$FeBr_3$ is “iron(III) bromide” (Fe^{3+} ion)
[speak “iron-3-bromide”]

Naming Compounds With <u>Polyatomic Ions</u>

1. Name the (+ve) ion first.

2. Add the name of the (-ve) ion. The name of a polyatomic ion does NOT change.

<u>Examples</u>

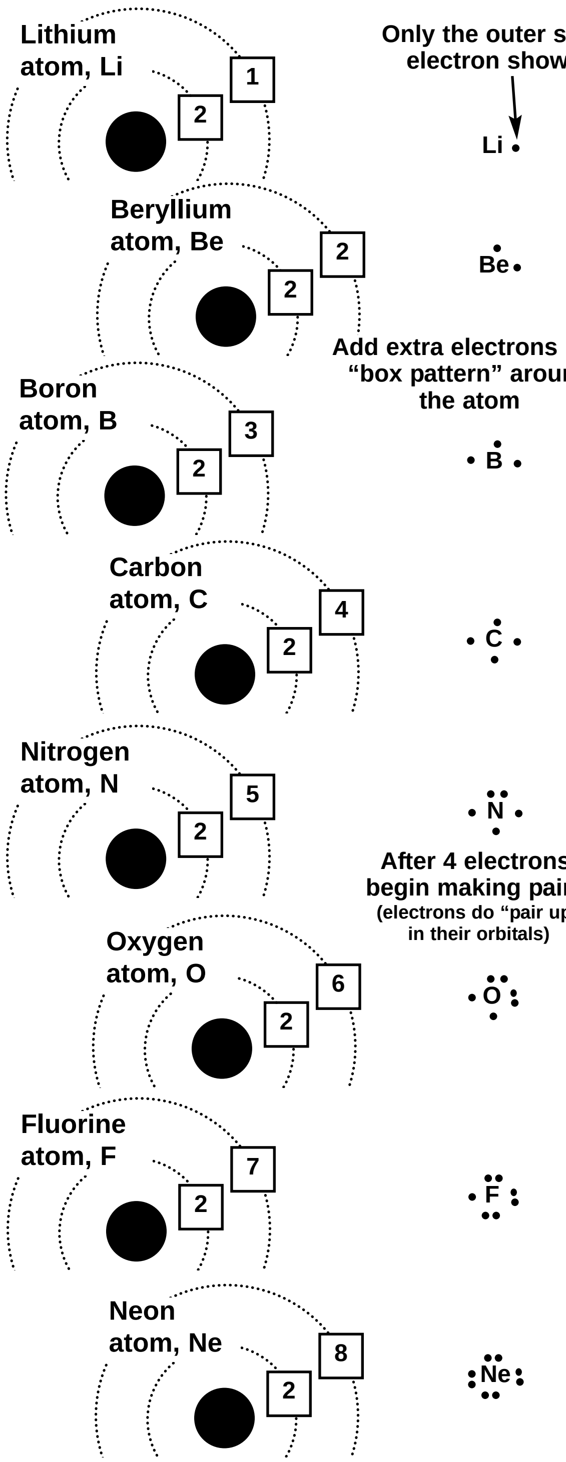
$Fe(NO_3)_2$ is “iron(II) nitrate”
$(NH_4)_2SO_4$ is “ammonium sulfate”

Lewis Formulas... "Electron-Dot Diagrams"

Before going any further you need to learn another way to represent atoms, ions and molecules and their electrons.

A "Lewis Formula" for a chemical species uses dots to represent the <u>outer shell electrons</u> of each atom. The rest of the atom is represented by the chemical symbol for that element.

<u>Examples</u>: **<u>Lewis Formulas</u>**



Only the outer shell electron shown

Add extra electrons in a "box pattern" around the atom

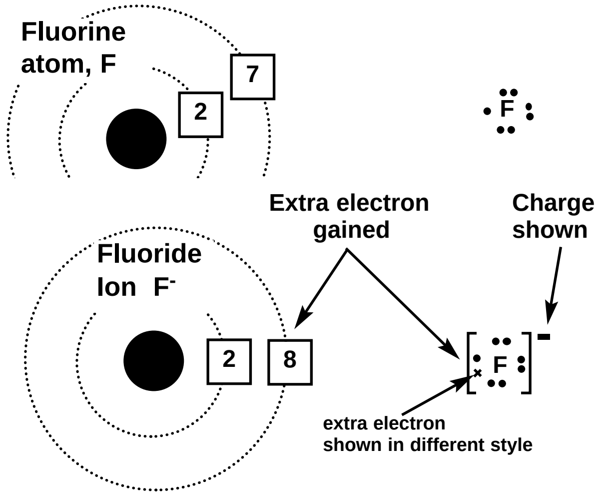
After 4 electrons, begin making pairs. (electrons do "pair up" in their orbitals)

Outer shell is complete with 8 electrons in 4 pairs. The next element (sodium) would have a Lewis Formula similar to lithium's.

Lewis Formulas for Ions

When a non-metal gains 1 or more electrons to form a negative ion, the extras are shown in a different style, for example:

<u>Lewis Formulas</u>

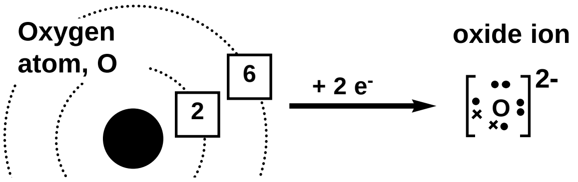


Extra electron gained

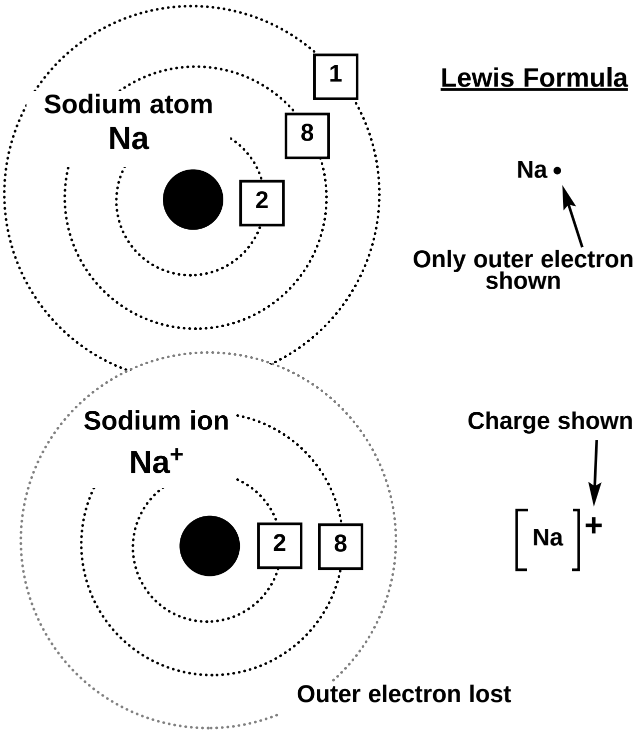
Charge shown

extra electron shown in different style

An oxygen atom gains 2 electrons to form the oxide ion:



A Lewis Formula is not very useful for showing simple positive ions, but for the record...



Only outer electron shown

Charge shown

Outer electron lost

Equations for Ion Formation

As well representing atoms and ions by various diagrams and formulas, you must learn that any change that occurs can be represented by an equation.

Formation of a Fluoride Ion

(from an atom):

$F + e^- \longrightarrow F^-$

Fluorine atom + extra electron → Fluoride ion

In fact, Fluorine exists in nature as F_2 molecules, not single atoms. If a molecule of F_2 formed ions, the equation is:

$F_2 + 2e^- \longrightarrow 2F^-$

Fluorine molecule + 2 extra electrons → 2 Fluoride ions

Formation of a Magnesium Ion

Note that in ALL cases the total amount of electric charge on each side of the equation is equal.

$Mg \longrightarrow Mg^{2+} + 2e^-$

Magnesium atom → Magnesium ion + 2 electrons

(Since Mg atoms have 2 electrons in their outer shell, they must lose them to form the normal ion with (2+) charge)

Formation of Magnesium Fluoride from its Elements

The last 2 equations can be simply added together to describe the formation of the ionic compound Magnesium fluoride (formula MgF_2) from its elements.

$F_2 + 2e^- \longrightarrow 2F^-$

add these together

$Mg \longrightarrow Mg^{2+} + 2e^-$

$Mg + F_2 + \cancel{2e^-} \longrightarrow Mg^{2+} + \cancel{2e^-} + 2F^-$

notice how 2 electrons occur on both sides, so they cancel out.

$Mg + F_2 \longrightarrow MgF_2$

These equations show how for each magnesium ion there needs to be 2 fluoride ions, so that everything “balances”.

Here is another example, but more complicated.

It demonstrates the importance of balancing chemical equations:

Formation of an Oxide Ion (from an atom)

$O + 2e^- \longrightarrow O^{2-}$

Oxygen atom + 2 extra electrons → Oxide ion

However, just like fluorine, oxygen always exists in nature as O_2 molecules, so the reaction would be:

Formation of Oxide Ions (from an O_2 molecule)

$O_2 + 4e^- \longrightarrow 2O^{2-}$

Oxygen molecule + 4 extra electrons → 2 Oxide ions

Where might the extra electrons have come from? In a chemical reaction, they would normally come from a metal atom which needs to lose electron(s) to achieve its best energy state.

Let’s assume the metal is Lithium:

Formation of a Lithium Ion

$Li \longrightarrow Li^+ + e^-$

Lithium atom → Lithium ion + electron

(Since Li atoms have 1 electron in their outer shell, they must lose it to form the ion with (+1) charge)

However, in a real situation where lithium is reacting with oxygen, each O_2 molecule needs 4 electrons. Therefore, it will take 4 lithium atoms to supply them...

Previous Equation Multiplied by 4

$4Li \longrightarrow 4Li^+ + 4e^-$

4 Lithium atoms → 4 Lithium ions + 4 electrons

Now add together the equations together:

$O_2 + 4e^- \longrightarrow 4O^{2-}$

add these together

$4Li \longrightarrow 4Li^+ + 4e^-$

$4Li + O_2 + \cancel{4e^-} \longrightarrow 4Li^+ + \cancel{4e^-} + 2O^{2-}$

There are 4 electrons on both sides, so they cancel out.

The correct formula for lithium oxide is Li_2O, so this combination of ions is enough to make “2 lots” of ions.

$4Li + O_2 \longrightarrow 2Li_2O$

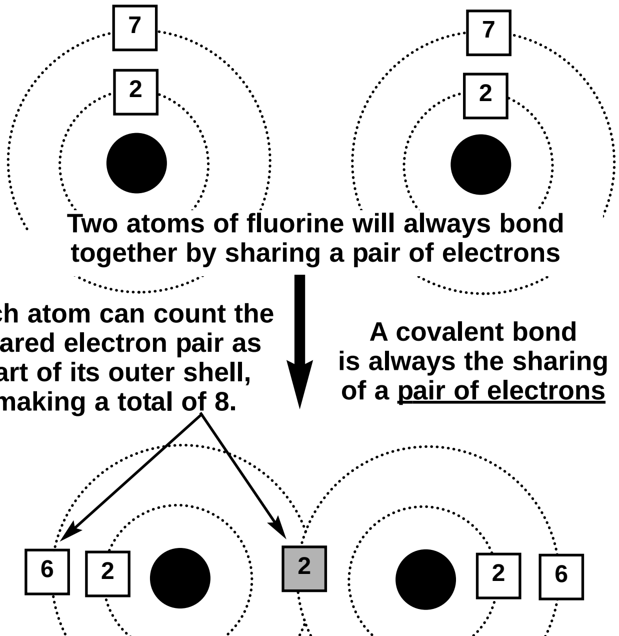
This equation is “balanced”... it shows the same number of each atom on both sides of the equation.

Covalent Bonding

The gaining and losing of electrons to form ions is not the only way for atoms to achieve a full outer shell, and the best possible energy state that goes with that. If atoms have 4, 5, 6 or 7 electrons in their outer orbit, they can also achieve a full outer shell by sharing electrons. This is called “covalent bonding”.

Not only do many compounds form this way, but many elements exist as 2 or more atoms covalently bonded together.

Example: the element Fluorine, F_2



Two atoms of fluorine will always bond together by sharing a pair of electrons

Each atom can count the shared electron pair as part of its outer shell, making a total of 8.

A covalent bond is always the sharing of a pair of electrons

One molecule of fluorine, F_2

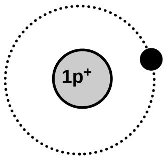
These atoms must remain tightly attached to each other in order to share the electrons... they are bonded together very strongly, forming a molecule of F_2.

Atoms of all the non-metals and the semi-metals, (except Group 8 the “Inert Gases”) bond together covalently when in the pure elemental state:

Elements	Molecules
Group 7	F_2 Cl_2 Br_2 I_2
Nitrogen, Oxygen	N_2 O_2
Sulfur	S_8
Phosphorus	P_4
Carbon, Silicon and others	covalent lattices

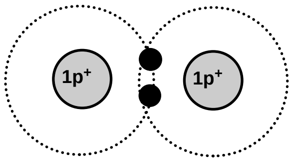
Hydrogen... the Weirdo

Hydrogen is the smallest atom of all, with only 1 electron.



Sometimes hydrogen loses this electron, forming a hydrogen ion H^+. When this happens it is behaving chemically like a metal in Group 1.

However, hydrogen atoms can also share electrons covalently. Elemental hydrogen is always H_2 molecules:



Don’t forget that the 1st orbit holds a maximum of 2 electrons, so both atoms achieve a full outer shell by sharing.

Hydrogen, and all the non-metals and semi-metals not only bond with atoms of the same type in the the element state, but will share electrons with different atoms to form covalent compounds.

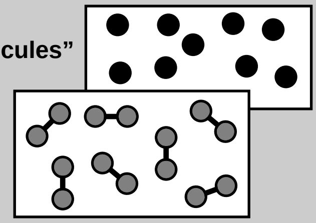
A Note About “Molecules”

You need to be aware of the precise definition of the word “molecule”.

Definition: A molecule is the smallest particle of a substance that can have a separate existence, and can move around independently of other particles.

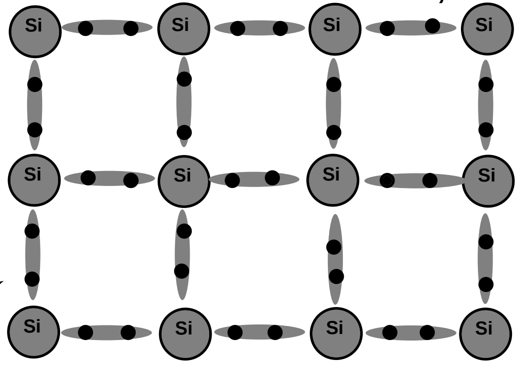
Examples:
Inert gases have “molecules” of just one atom.

Hydrogen has “diatomic” molecules (di=2)



Lattice structures (ionic or covalent) are not molecules.

A Covalent Lattice Element; Silicon



Each chemical bond is formed by atoms sharing 2 electrons with each neighbour

The shape of the lattice is not square as in this 2-dimentional diagram. In 3-dimensions, the atoms surround each other in a pyramid pattern, called a “tetrahedron”.

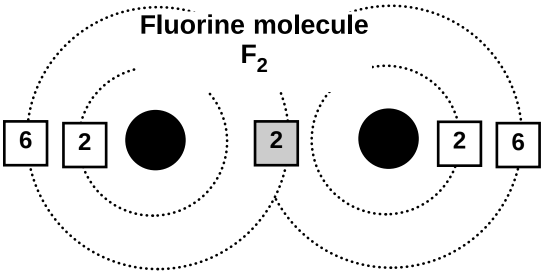
Pure silicon forms hard crystals made of billions of atoms covalently bonded in a lattice.

Lewis Formulas with Covalent Bonding

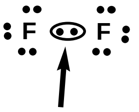
**You were introduced to the Lewis Formula ("electron-dot" diagram) previously.
Lewis Formulas are actually more useful for covalent situations than they are for ions.**

A covalent bond is always a pair of electrons being shared between 2 atoms. In a Lewis Formula the shared pair are usually emphasised by drawing a little oval around them.

Example:



Lewis Formula

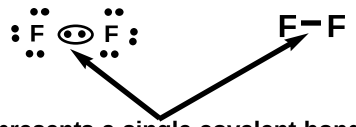


Shared pair of electrons

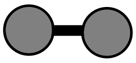
The F_2 Fluorine molecule could also be represented by a simple structural formula showing that it contains 2 atoms of fluorine which are connected by a single covalent bond.

Molecular Formula | **Lewis Formula** | **Structural Formula**

F_2



**This represents a single covalent bond
(Remember this involves sharing 2 electrons)**



Model of the molecule

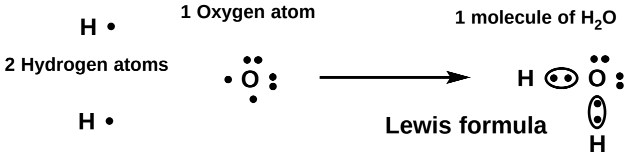
You need to be able to interpret, and construct, all of these different ways to represent atoms and molecules.

Covalent Compounds

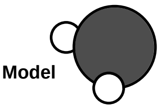
Many common and important substances are formed by covalent bonding between atoms of 2 or more different elements.

Understanding Covalent Compounds with Lewis Formulas

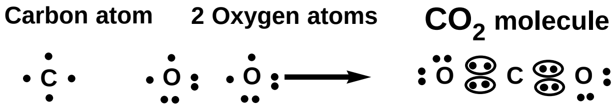
Everyone knows that water is H_2O. You need to understand exactly how this compound forms.



Look carefully at the Lewis formula above to see how all the atoms involved have achieved full outer shells of electrons by sharing pairs in covalent bonds.



Another well known covalent molecular compound is carbon dioxide CO_2



The CO_2 molecule contains double covalent bonds. These involve atoms sharing 2 pairs of electrons. The structural formula for this would be:

O=C=O

It's also possible to have a triple covalent bond; 3 pairs = 6 electrons being shared between 2 atoms. This occurs in the nitrogen (N_2) molecule below, as well as other compounds.

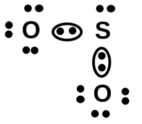
N≡N

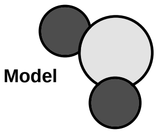
Predicting Formulas for Covalent Compounds

The formulas of the examples at left are quite predictable if you know how many electrons are in each atom's outer shell, and understand how sharing electrons can achieve a full outer shell.

However, not all covalent compounds are so predictable, because the "rule" about achieving a full shell of 8 electrons is not always followed with covalent bonding. (It is always followed with ionic bonding.)

For example, if oxygen & sulfur combine covalently, the compound formed is sulfur dioxide (SO_2).





Study this Lewis Formula and you'll see that the "rule of 8 electrons" has NOT been followed for the sulfur atom!

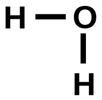
Naming Covalent Compounds

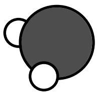
The first problem you face here is that (for historical reasons) many covalent compounds have “common names” that follow no rule or system, and must be memorised.

Common Names

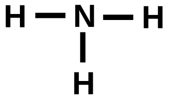
To keep this as simple as possible (K.I.S.S. Principle!) start with just these three common, important compounds:

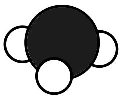
Water H_2O





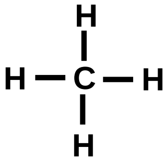
Ammonia NH_3

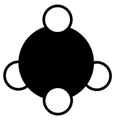




(Not to be confused with the ammonium polyatomic ion (NH_4^+))

Methane CH_4





(This is the simplest of a huge range of covalent compounds of carbon & hydrogen... more in later topics)

More Than One Compound

The second problem is that, quite often, there is more than one possible compound formed from the same elements in a covalent compound.

Some examples

Different Compounds Possible			***Different Compounds Possible***		
Sulfur & oxygen	SO_2 Sulfur dioxide	SO_3 Sulfur trioxide	**Carbon & oxygen**	CO Carbon monoxide	CO_2 Carbon dioxide

To cope with this, a naming system has developed which uses prefixes to state how many atoms of each element are in one molecule.

The Prefixes

1 = mono 2 = di 3 = tri 4 = tetra 5 = penta 6 = hexa

How to Name a Simple Covalent Compound From a Molecular Formula

- **Name the elements in the order as in the formula.**
- **Alter the name of the 2nd element to end -IDE.**
- **Attach a prefix to both names, to show how many atoms are in each molecule.**

(Important exception: If there is only one atom of the first-named element, do not attach a prefix to it.)

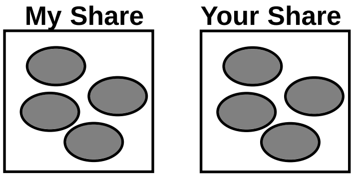
Examples: **P_2O_5 = diphosphorus pentoxide** **SO_3 = sulfur trioxide**

N_2S_3 = dinitrogen trisulfide

Ionic v. Covalent & Bonds In-Between

Up to this point, you have seen ionic & covalent bonding as quite different things. Now you must realise that they are just different degrees of the same thing.

An analogy might help... Imagine 2 people sharing some lollies. If both people are very fair about it, and neither dominates or intimidates the other, the sharing will be equal:

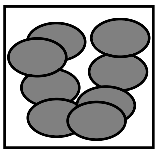


This is like a "pure <u>covalent bond</u>" where electrons are shared equally

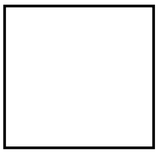
In contrast, an <u>ionic bond</u> can be thought of as the lolly-sharing between a hungry bully and a wimp who hates lollies:

<u>Bully's Share</u>
"Gimme everything"

<u>Wimp's Share</u>
"I don't want any"

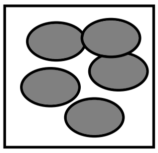


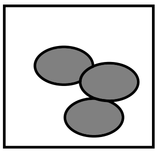
When electrons are shared so unequally, the result is (+ve) and (-ve) ions being formed.



Now you must learn that there is also a situation (or a whole heap of situations) in between these extremes, where the lollies will be shared, but perhaps not evenly.

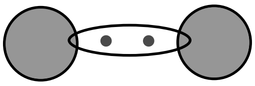
Sharing, but not equally





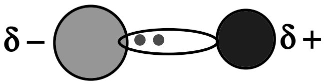
In chemical bonding, this kind of sharing is called a "Polar Covalent Bond" and occurs when electrons are shared between 2 atoms with quite different values for <u>Electronegativity</u>.

Polar Covalent Bonds



A "Pure" Covalent Bond occurs when electrons are shared evenly.

In a "Polar Covalent Bond" the sharing is not even. The electrons are attracted more to one atom than the other.



This is a "dipole"
It has 2 opposite poles

This causes the bond (and the entire molecule) to become electrically "polarised". The electric charge is not evenly distributed. One end has a greater concentration of electrons and has a slight negative charge (δ–), while the other end becomes slightly positive (δ+).

The Greek letter delta (δ) is used to denote a "small amount" of something, in this case electric charge. The molecule is called a "dipole", meaning it has 2 poles.

Electronegativity Difference

You were introduced to the concept of electronegativity earlier in this topic... revise if necessary.

The higher the value of electronegativity for any element, the more it attracts electrons to its atomic nucleus. The highest value elements are fluorine, oxygen, nitrogen & chlorine, clustered in the top right corner of the Periodic Table. (But Inert Gases don't count)

The lowest values are among the metals of groups 1 & 2, especially lower down the table.

When any 2 atoms react & form a chemical bond, it is the <u>difference in electronegativity</u> which determines the nature of the bond.

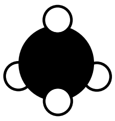
Electronegativity & Bond Type

Pure Covalent Bonds

If the difference in electronegativity is small (less than about 0.5 units) the bond is considered <u>purely covalent</u>.

Examples:
If 2 identical atoms bond together (eg F_2, O_2, H_2, Cl_2, etc.) there is NO difference in electronegativity... pure covalent.

When hydrogen bonds to carbon (thousands of examples) the difference is only 0.4 units... pure covalent.

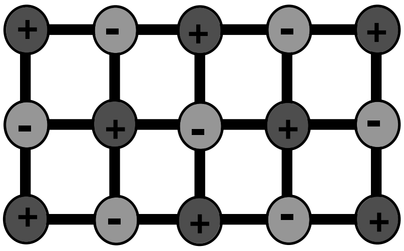


methane CH_4

Ionic Bonds

If the the difference is more than about 1.4 units, the electrons are attracted to one atom so much more than the other that the bond is considered <u>ionic</u>. It is as if the electrons have been totally transferred from one to the other.

Examples:
Almost any metal combining with chlorine, oxygen or other non-metals will result in a typical ionic compound. In the solid state these form a "crystal lattice" of ions.

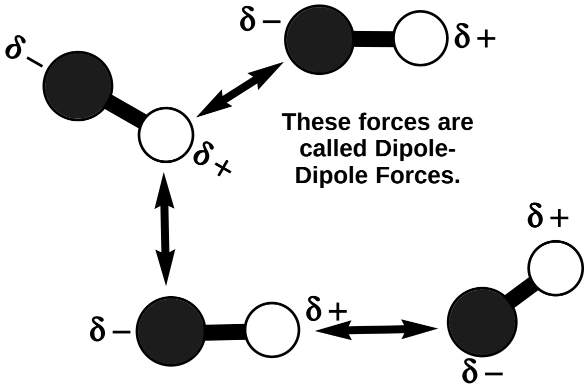


Polar Covalent Bonds

If the difference in electronegativity is <u>more than 0.5, but less than 1.4 units</u>, the bond is likely to be <u>polar covalent</u>. The compound will probably form discrete molecules, but they may exhibit "polarity" because parts of the molecule will have slight electrical imbalance. (It also depends on the exact shape of the molecule. Perfectly symmetrical molecules can have polar bonds within, but the symmetry might cancel out any uneven electrical charges.)

The consequences of polarity are explained on the next page.

Polar Bonds Create Inter-Molecular Forces



The charges on each end of a molecular dipole are only a fraction of the size of the charges on an ion, but they do cause electrical forces to occur between nearby molecules.

It is these forces which are the “<u>inter-molecular forces</u>” that make the molecules cling together in the solid state. These are the forces which must be overcome with heat energy to melt the solid. These are the forces which determine the m.p. and b.p. of a molecular substance.

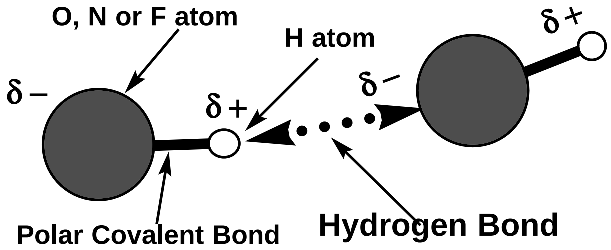
The strength of the dipole-dipole force varies according to the degree of polarity of the covalent bond (how evenly or unevenly the electrons are being shared) and also varies according to the shape of the molecule. In some substances the forces are quite weak, in others quite strong.

The strongest dipole-dipole forces are about 1/3 as strong as a full-scale ionic bond. These occur whenever <u>hydrogen</u> atoms are bonded to <u>oxygen,</u> <u>nitrogen</u> or <u>fluorine</u>, and are called...

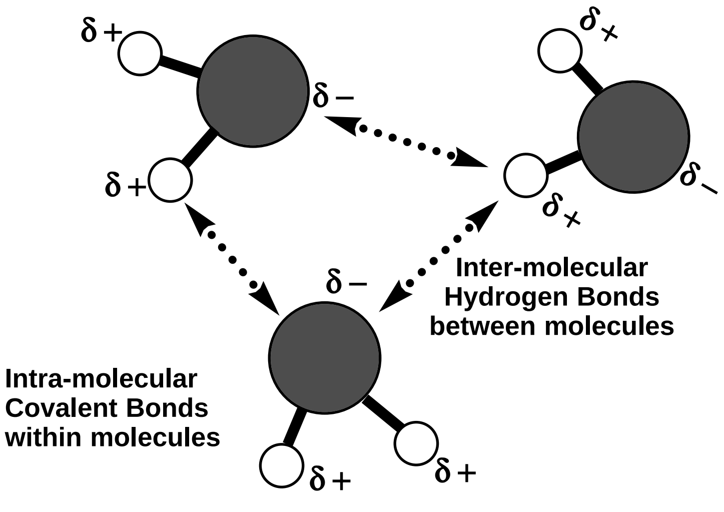
Hydrogen Bonds

Oxygen, Nitrogen and Fluorine are all small, strongly electronegative atoms. Hydrogen is even smaller, and once the electrons are “sucked away” from it in the polar bond, the hydrogen atom is really a “naked” proton.

The result is an especially strong set of partial charges, a powerful dipole, and strong inter-molecular force, which attracts nearby molecules to each other. These especially strong dipole-dipole attractions are called “Hydrogen Bonds”.



Hydrogen Bonding in Water



In the water molecule the covalent bonds are very polar, so the atoms develop especially large partial charges. Each molecule is a dipole, and strong inter-molecular “Hydrogen Bonds” attracts each molecule to its neighbours.

It is this network of hydrogen bonds that holds the molecules in a rigid lattice in the solid state. (ice)

The Hydrogen Bonding is the reason that ice has such a high melting point, compared to other molecules of similar size. (Ammonia also has relatively high m.p. & b.p... same reason!)

Once melted to a liquid, the molecules can move around, but “cling” to each other because of the hydrogen bonds. The molecules even “wriggle” closer to each other and the density increases.

To boil water to a gas, the molecules must be able to totally break free from the hydrogen bonds. This requires considerable energy, so water has an unusually high boiling point, compared to other silimar sized molecules.

It is the HYDROGEN BONDING between water molecules which explains water’s unusual properties:

- **Water has much <u>higher mp & bp</u> than other similar sized molecules.**
- **Water is one of very few substances in which the liquid is more dense than the solid. (That’s why ice floats in water.)**
- **Water has very <u>high viscosity</u> for such a small molecule. (Viscosity is a measure of how “sticky” water is.)**
- **Water has an unusually strong “<u>surface tension</u>”. (A network of forces at the surface which can support things which should sink.)**

5. Bonding, Structures & Properties

Physical and Chemical Properties

How do you recognise things and tell them apart?

How, for example, do you tell an orange from a banana? You look at its colour and shape and (if blind-folded) you'll go by smell and taste. You are using the properties of different things to identify them.

In Chemistry it's exactly the same... we identify substances, and classify different chemicals according to their properties. What are the properties we use?

Physical Properties

Melting & Boiling Points. (these determine the state)
Electrical Conductivity.
"Hardness" and Flexibility. (e.g. malleability and ductility)

There are many other properties, such as density and colour, but the three above are by far the most useful when surveying and classifying matter in a general way (and using the K.I.S.S Principle!).

Chemical Properties including:

- how reactive the substance is.
- whether it is acidic, basic or neutral.
- which types of reactions it will undergo. (e.g. whether it will burn or corrode)

Chemical properties are not so important when surveying and classifying matter in a general way, but will become important in later topics.

Properties of Elements, Compounds & Mixtures

Elements and Compounds are all pure substances. Each element, and each compound has its own unique properties which are characteristic and do NOT vary.

For example, pure water has a fixed melting point, boiling point, density, acidity, conductivity, etc. It is these unique, fixed properties which allow us to recognise and identify water, and every other pure substance.

Mixtures are not pure. The properties of mixtures are usually a "blend" of the properties of its parts, and vary according to its exact composition. For example, salt water has properties of both water and salt, and its density, boiling point, conductivity (and taste) vary according to the proportions of the mixture.

**Elements & Compounds are Pure Substances
with Unique, Fixed Properties.
Mixtures are Impure
and have "blended", Variable Properties**

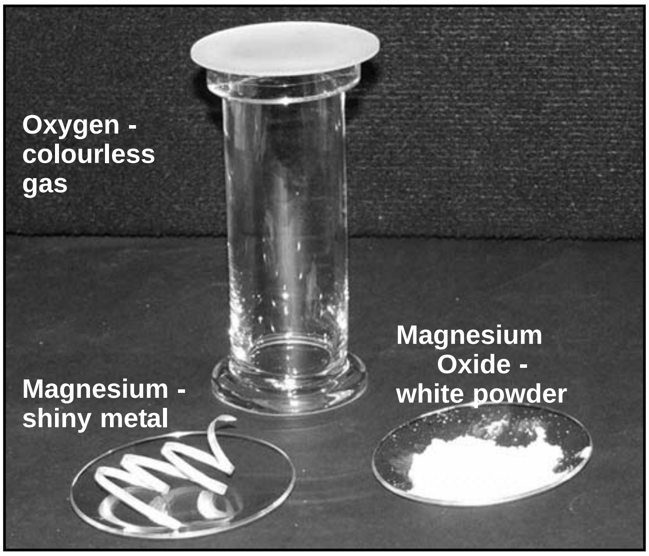
Despite mentioning "taste" several times on this page, tasting is NOT safe or appropriate in the laboratory. Don't taste the chemicals!!

Properties of a Compound Compared to Its Elements

You may have done practical work to investigate whether or not the properties of a compound are related to the properties of the elements it contains.

A simple example is to examine a piece of magnesium (element) and note some basic physical properties. Then consider the observable properties of the element oxygen, in the air around us.

Then burn the magnesium in air. The product of the reaction is the compound magnesium oxide, which can be collected and its properties noted.



Magnesium: metallic solid... shiny, flexible, conductor.
Oxygen: colourless, odourless, non-conducting gas.

Magnesium oxide: brittle, powdery, white, non-conducting solid.

Consider the compound sucrose (table sugar) and the elements carbon, hydrogen and oxygen it is made from.

Carbon: black, brittle solid.
Hydrogen: colourless, explosive gas.
Oxygen: colourless, odourless gas

Sucrose: clear, crystalline solid, with a sweet taste.

You may have examined and considered many other examples. The general conclusion is:

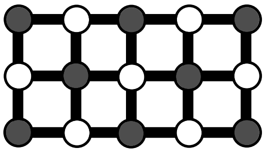
Generally, the properties of a Compound are totally different to the properties of its Elements

Bonding Within Substances

To survey and understand the general categories of matter, it is important to know the different forces that operate to hold substances together. It is this "bonding" within substances that often determines the general physical properties by which we classify matter into types.

Ionic Compounds

are a lattice of (+ve) and (-ve) ions.



The "Ionic Bonds" are actually electrical attractions between opposite charges, and are very strong.

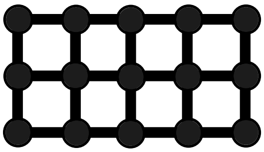
Since these bonds are strong, it requires a lot of energy to break them so that the particles can begin moving around.
Therefore, the melting and boiling point is usually quite high.

The substance usually has a hard, crystalline structure because all the ions are locked into a 3-D "lattice". The crystal shape reflects the arrangement of ions in the lattice.

Covalent Lattice Substances

Some elements (e.g. Carbon and Silicon) and some covalent compounds (e.g. silicon dioxide, SiO_2) form a lattice of atoms covalently bonded, in a 3-dimensional crystal structure.

Silicon dioxide forms the mineral "silica", the most common on Earth. A grain of sand is a crystal of silica.



The "Covalent Bonds" are formed when a pair of electrons is being shared, and are very strong.

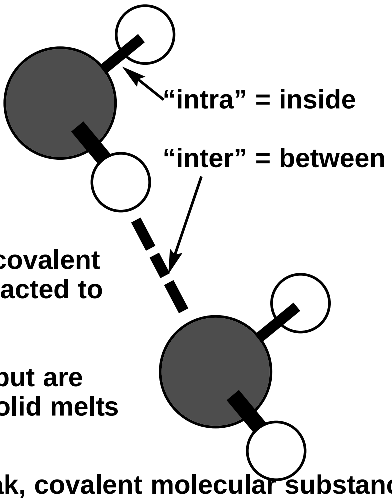
Since these bonds are strong, it requires a lot of energy to break them so that the particles can begin moving around. Therefore, the melting and boiling point is usually very high.

Like the ionic lattices, the covalent lattice substances may form shiny, hard crystals. The big difference between them is that ionic lattices are often soluble in water; covalent lattices never are.

Covalent Molecular Substances

Some elements (e.g. oxygen, chlorine) and many compounds (e.g. water, carbon dioxide) are composed of covalent molecules.

To understand these substances you must know about "intra-molecular" and "inter-molecular" forces.



The forces between the molecules are due to polar covalent bonds creating "dipoles". Nearby molecules are attracted to each other by the opposite electrical charges.

They hold the molecules in place in the solid state, but are easily broken by heat energy. This means that the solid melts easily.

The Covalent Bonds inside the molecules are very strong and require a lot of energy to break.

It may take a lot of energy to decompose the compound.

The bonds between molecules are weaker, so mp's & bp's are low compared to ionic or covalent lattices.

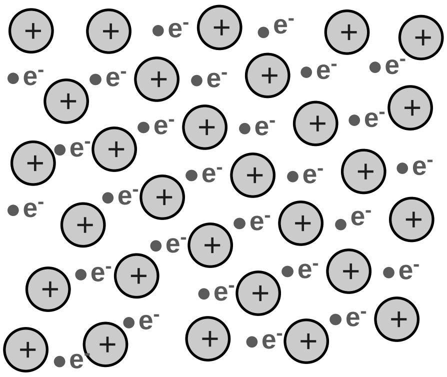
Since the "inter-molecular" forces are relatively weak, covalent molecular substances generally have relatively low melting and boiling points, and many are liquids or gases at standard temperature & pressure.

Bonding Within Metals

Why are most metals hard, with quite high melting points? There must be some strong bond holding the atoms together, yet allowing them to change shape (malleable & ductile) when hammered or stretched.

Metal atoms do not hold onto their outer (valence) electrons. Each atom is actually a (+ve) ion. Loose electrons wander between the ions, in a shifting "sea of electrons".

The "Metallic Bond" is the electrical attraction between the (+ve) ions and the surrounding "sea" of negative charges.

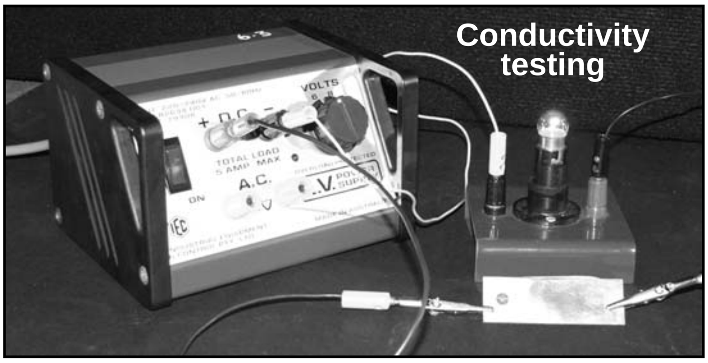


This bond can be very strong in some metals, so the metal is strong, with a high melting point. In other metals the bond is weaker, so some metals are softer and melt at lower temperatures. (e.g. lead)

However, while the metallic bond can be very strong, it is not rigid. The sea of electrons shifts and flows, so the ions can be pushed or pulled to different places without breaking the substance. This is why metals are malleable and ductile, unlike the hard, but brittle ionic or covalent lattices.

Comparing the Properties of Different Substances

You may have done practical work to study the properties of a variety of substances.



The properties studied were probably:

- mp & bp (from Chemical Data book or table)
- electrical conductivity, in solid & liquid states, and in solution (if soluble) by experiment.
- hardness and flexibility of the solid, by experiment.

Typical General Results

Category	Melting Pt (°C)	Boiling Pt (°C)	Electrical Conductivity Solid	Liquid	Solution	Hardness/Flexibility of solid
<u>Metals</u> (e.g. Iron, Lead)	Medium to High	High	Good	Good	N/A	Most hard, malleable & ductile
<u>Ionic Compounds</u> (e.g. Salt NaCl Sodium hydroxide)	Medium to High	High	Poor	Good	Good	Hard & brittle
<u>Covalent Lattices</u> (e.g. Silicon dioxide diamond (carbon)	Very High	Very High	Poor* (Si, Ge are semi-conductors)	Poor	N/A	Hard* & brittle
<u>Covalent Molecules</u> (e.g. water, carbon dioxide)	Low	Low to medium	Poor	Poor	Poor	Solids often soft & waxy. If hard, then brittle. (e.g. water ice)

<u>Exceptions & Anomalies</u>:
* Carbon, in the form of graphite, is a good conductor, and is soft and slippery.

Explaining Electrical Conductivity

Any substance will conduct electricity if it contains electrically charged particles which can move independently of each other.

Metals contain metal ions and a mobile "sea" of free electrons. When a voltage is applied, electrical current is carried readily by the electrons flowing among the metal ions.

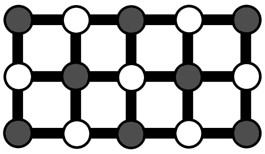
This makes metals <u>good conductors</u> in both solid and liquid states.

Covalent Lattices & Covalent Molecules do NOT contain any charged particles that can separate from each other and move independently..

These substances are generally poor conductors whether solid, liquid or in water solution.

(<u>Exceptions</u>: Graphite is a good conductor. The "semi-metal" elements (notably Si & Ge) are semi-conductors)

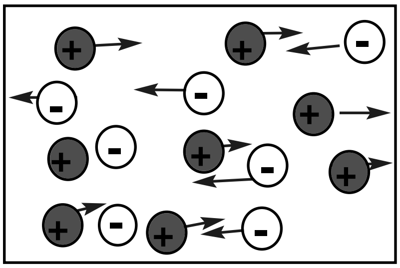
Ionic Compounds are the trickiest to understand! In the solid state the ions are fixed in the lattice and cannot move freely.



Solid ionic lattices will NOT conduct because ions cannot move freely.

However, if an ionic compound is melted, the ions can move freely and independently. If a voltage is applied, a current will be carried by the ions migrating in opposite directions.

Many ionic compounds are soluble in water. When they dissolve, the lattice disintegrates and the ions can move freely. (This will be explained fully in a later topic)



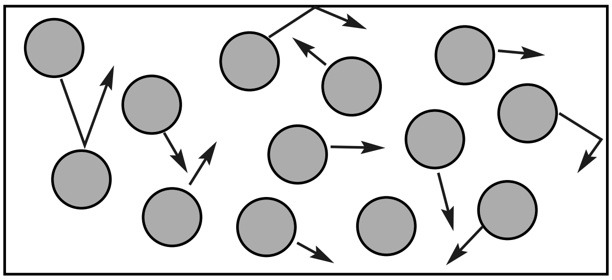
Ionic compounds become good conductors in the liquid state, and in solution.

A Few Loose Ends to Tidy Up

A Note About the Inert Gases

What about the elements of Group 18, the "Inert Gases"?

These elements have full outer shells of electrons, so they do not normally form ions, nor share electrons covalently. Therefore, they always exist as <u>single-atom molecules</u>. (Remember the exact definition of a "molecule") Technically, therefore, they are molecular substances. When we write "He" for helium this is both the atomic symbol and the molecular formula.



In this case there are no covalent bonds within molecules. There are, however, some <u>extremely weak</u> inter-molecular forces which can hold the atoms in a solid lattice at extremely low temperatures. Even very tiny amounts of heat can overpower these forces, so helium melts and boils to a gas at a temperature around -270°C.

The forces get stronger as the atoms get bigger, but even so, all the elements of Group 8 are gases at room temperature because of very low m.p.'s and b.p.'s.

Molecular and Empirical Formulas

When we say that the formula for water is H_2O, we mean that each molecule of water contains 2 atoms of hydrogen and 1 atom of oxygen. "H_2O" is a <u>molecular formula</u> which describes the molecules.

Salt is an ionic compound. Each crystal contains billions of sodium and chloride ions, but they are in the ratio of 1:1. The formula is NaCl, which is an <u>empirical formula</u>. It does not describe molecules (there aren't any!) but gives the simplest ratio of the elements present.

Similarly, <u>silicon dioxide</u> has the formula SiO_2, but there are no molecules. This compound is a covalent lattice of billions of atoms bonded together. The atoms are in the ratio of 1 silicon atom to every 2 oxygen atoms. SiO_2 is an empirical formula.

One Last Thing... Allotropes

You learnt earlier about "isotopes" of the elements. Well, just to confuse you, the syllabus wants you to know about "allotropes" as well. Allotropes are different forms of the same element with <u>different physical properties</u> due to having different molecular or lattice structures.

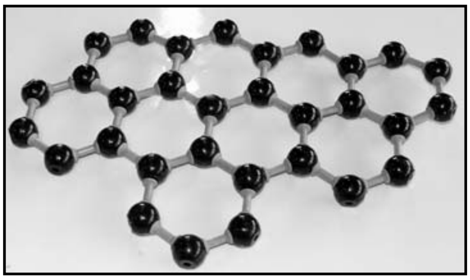
Allotropes are composed of the same atoms, but arranged differently. This gives them different densities, mp & bp, colour & appearance, conductivity, etc. Perhaps the best known allotropes are those of the element carbon. Excellent for a case study:

...next page

Case Study: Allotropes of Carbon

Graphite

In Graphite, the carbon atoms are arranged in hexagonal rings which connect to form flat sheets.

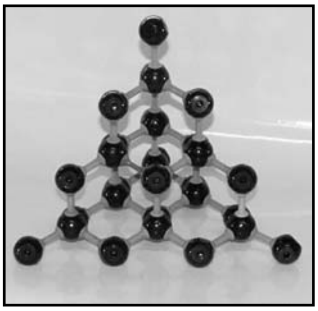


The atoms within each sheet are strongly bonded, (m.p. & b.p. are high) but the bonds between the sheets are very weak. They can easily slide past each other, so graphite is "slippery". Because of its "slipperiness", graphite is an excellent lubricant, used for example, in door locks. Its most familiar use is the "lead" in a pencil.

Unusually for a non-metal and covalent lattice, graphite is a good conductor of electricity.

Diamond

is also a covalent lattice of carbon atoms, but the atoms are arranged in a tetrahedral pattern, forming a huge 3-D crystal lattice.



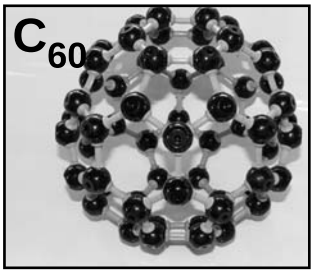
Diamond has extremely high m.p. & b.p., and is the hardest natural substance known.

The beauty of its sparkling crystals has made diamond treasured (literally) for jewellery throughout history.

In modern times, its extreme hardness has resulted in the widespread use of "industrial diamonds" for drill bits (e.g. for oil-drilling equipment to bore through solid rock) and high-speed saw blades.

"Bucky-Balls"

Discovered much more recently is a variety of allotropes of carbon... the Buckminster-Fullerenes. Named after the architect who invented the "geodesic dome" structure, Bucky-Balls and Bucky-Tubes come in a variety of shapes and sizes.



The best known has the formula C_{60}, in which the carbon atoms are arranged to form a sphere resembling a soccer ball.

The Bucky Balls have not yet found a practical use, but they have potential for use as high temperature lubricants, for making super-conducting polymers or even as specialised "capsules" for administering medicines.

ALLOTROPES = Same element, same atoms.
Different atomic arrangements, different physical properties.
Some other elements which have allotropes include
sulfur, phosphorus, oxygen and tin.

VALENCY TABLE
of common ions

Name	Symbol	Electrons in outer shell	Charge on Ion
Simple Metal Ions			
Hydrogen	H	1	1+
Lithium	Li	1	1+
Sodium	Na	1	1+
Potassium	K	1	1+
● Silver	Ag	1	1+
Magnesium	Mg	2	2+
Calcium	Ca	2	2+
Barium	Ba	2	2+
● Zinc	Zn	2	2+
Aluminium	Al	3	3+
Metals with More Than One Ion			
● Copper	Cu	1 or 2	1+ or 2+
● Iron	Fe	2 or 3	2+ or 3+
● Tin	Sn	2 or 4	2+ or 4+
● Lead	Pb	2 or 4	2+ or 4+
Polyatomic Ion			
● Ammonium	NH_4^+		1+

Name	Symbol	Electrons in outer shell	Charge on Ion
Simple Non-Metal Ions (name changes to -IDE)			
Fluorine	F	7	1-
Chlorine	Cl	7	1-
Bromine	Br	7	1-
Iodine	I	7	1-
Oxygen	O	6	2-
Sulfur	S	6	2-
Nitrogen	N	5	3-
Phosphorus	P	5	3-
Carbon	C	4	4-
Polyatomic Ions			
● Hydroxide	OH^-		1-
● Nitrate	NO_3^-		1-
● Sulfate	SO_4^{2-}		2-
● Carbonate	CO_3^{2-}		2-

Species marked “●” need to be memorised.
All others can be easily read from the Periodic Table.
As you encounter new ions (especially polyatomic), add to this table