

PiXL KnowIT!

GCSE Chemistry

AQA Topic – Atomic structure and the periodic table

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Atomic structure and the periodic table

Atoms, elements, compounds and mixtures

- Atoms, elements and compounds
- Word and symbol equations
- Separation techniques

Atomic model

- History of the atomic model
- Size and mass of atoms
- Atomic structure

Periodic table

- History of the periodic table
- Group 0, group 1 and group 7
- *Transition metals (chemistry only)*



All substances are made of **atoms** that cannot be chemically broken down. It is the smallest part of an **element**.

Elements are made of only one type of atom. Each element has its own **symbol**.
e.g. Na is sodium.

There are about **100** different elements

[illegible]

Compounds contain more than one type of atom. They are formed from elements by **chemical reactions**, which always involve the formation of one or more new substance, often involving an **energy change**.

The components of a compound **cannot** be **separated** by **physical** means. They can be separated only by **chemical** means.

You need to:

- Know the names and symbols for the first 20 elements and all of group 1, 7 and 0
- Name compounds when given the formulae or symbol equations



Compounds contain **two or more** elements **chemically** combined in **fixed proportions** and can be represented by formulae using the symbols of atoms from which they formed.

Naming a compound with **two elements** (usually a metal and a non metal) apply these rules:

- The metal name does not change
- The non-metal's name ends in **ide**

Naming a compound with a metal that reacts with ions that consist of **two or more** non-metal atoms covalently bonded together apply these rules:

- The metal name does not change
- The non-metal's name ends in **ate** if oxygen is present

For example:

Na_2S = sodium sulfide

K_2O = potassium oxide

When a compound contains a transition metal, the names become a bit more complicated. To distinguish them, Roman numerals indicate the charge on the metal ion

e.g. iron(II) chloride.

For example: Na_2CO_3 = sodium carbonate
 KNO_3 = potassium nitrate

Chemical equations: Show the **reactants** (what we start with) and the **products** (what we end with).

No **atoms** are **lost** or **made**. The **mass** of the **products** equals the **mass** of the **reactants**.

You need to:

- Write word equations
- Balanced symbol equations

Word equation: methane + oxygen ☐ carbon dioxide + water

Symbol equation: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$

There are 4 hydrogens here, bonded together.

There are 2 molecules of oxygen not bonded together.

There are 4 hydrogens here. You multiply the big number by the little number.



Equations **MUST** balance

- We can **ONLY** add **BIG** numbers to the front of a substance
- We can tell elements within a compound by **BIG** letters
- We can check an equation is balanced by counting the number of each type of atom on either side

We can add **state symbols** to a symbol equation to show whether the reactants and products in a chemical reaction are solids, gases, liquids or dissolved in water.

Solid = s

Liquid = l

Gas = g

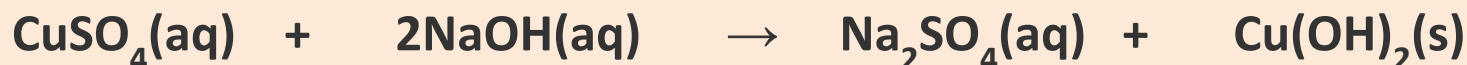
Aqueous (dissolved in water) = aq

Word equation: sodium + water \rightarrow sodium hydroxide + hydrogen

Symbol equation: $2\text{Na}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{NaOH}(\text{aq}) + \text{H}_2(\text{g})$

The reaction between copper sulfate and sodium hydroxide is:

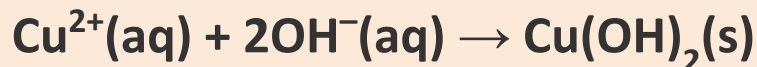
copper sulfate + sodium hydroxide \rightarrow sodium sulfate + copper hydroxide



You can tell that the copper hydroxide forms a solid (the precipitate) because its state symbol is (s) for solid, rather than (aq) for aqueous (dissolved in water).

HT only – write balanced half equations and ionic equations

The reaction can also be shown by an **ionic equation**:



QuestionIT

Atoms, elements, compounds and mixtures PART 1

- Atoms, elements and compounds
- Word and symbol equations



1. Around how many different types of element are there in the periodic table?
2. What does it mean if a compound ends in –ide?
3. What does it mean if a compound ends in –ate?
4. What does a chemical reaction involve?
5. Name the following substances:
 - NaCl
 - NaSO₄

6. Why is it useful to have symbols for atoms of different elements?
7. What is the difference between an element and a compound?
8. Explain why it is difficult to separate a compound, compared to an element.

9. Solid sodium reacts with water to form a sodium hydroxide (NaOH) solution and hydrogen gas.

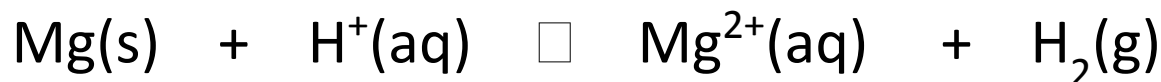
a) Write a word equation to represent this reaction.

b) Give the balance symbol equation for the reaction.

HT only:

10. When magnesium reacts with sulfuric acid, the hydrogen ions in the acid will be displaced from the solution by magnesium.

Balance the following ionic equation.



AnswerIT!

Atoms, elements, compounds and mixtures PART 1

- Atoms, elements and compounds
- Word and symbol equations



1. Around how many different types of elements are there in the periodic table?

1000

2. What does it mean if a compound ends in –ide?

Contains 2 elements; usually metal and non-metal.

3. What does it mean if a compound ends in –ate?

Contains 3 or more elements, 1 of which is always oxygen.

4. What does a chemical reaction involve?

The formation of one or more new substances; usually with an energy change.

5. Name the following substances:

- NaCl **Sodium chloride.**
- NaSO₄ **Sodium sulfate.**

6. Why is it useful to have symbols for atoms of different elements?
When elements join together to form a compound, it tells you how many atoms there are.

7. What is the difference between an element and a compound?
Elements contain one type of atom and compounds contain more than one type.

8. Why is it difficult to separate a compound?
A compound's elements are chemically joined together.

9. Solid sodium reacts with water to form a sodium hydroxide solution and hydrogen gas.

a) Write a word equation to represent this reaction.

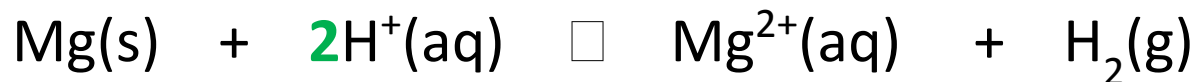
Sodium + water \rightarrow sodium hydroxide + hydrogen

b) Give the balance symbol equation for the reaction.

$2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + \text{H}_2\text{(g)}$

HT only:

10. When magnesium reacts with sulfuric acid, the hydrogen ions in the acid will be displaced from the solution by magnesium.
Balance the following ionic equation.



- **Mixtures**

A mixture consists of **two or more** elements or compounds **not** chemically combined together. The chemical properties of each substance in the mixture are **unchanged**.

Solvent	the liquid in which a solute dissolves
Solute	the substance that dissolves in a liquid to form a solution
Solution	is the mixture formed when a solute has dissolved in a solvent
Soluble	describes a substance that will dissolve
Insoluble	describes a substance that will not dissolve

FILTRATION: This technique separates substances that are insoluble in a solvent from those that are soluble

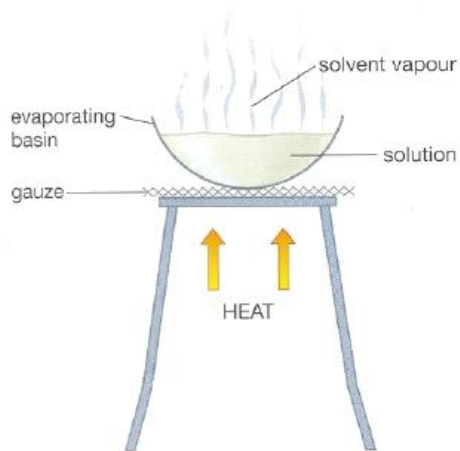
Mixtures can be separated by **physical processes** including:

1. Filtration
2. Crystallisation
3. Simple distillation
4. Fractional distillation
5. Chromatography

These physical processes do not involve chemical reactions and no new substances are made.

Crystallisation

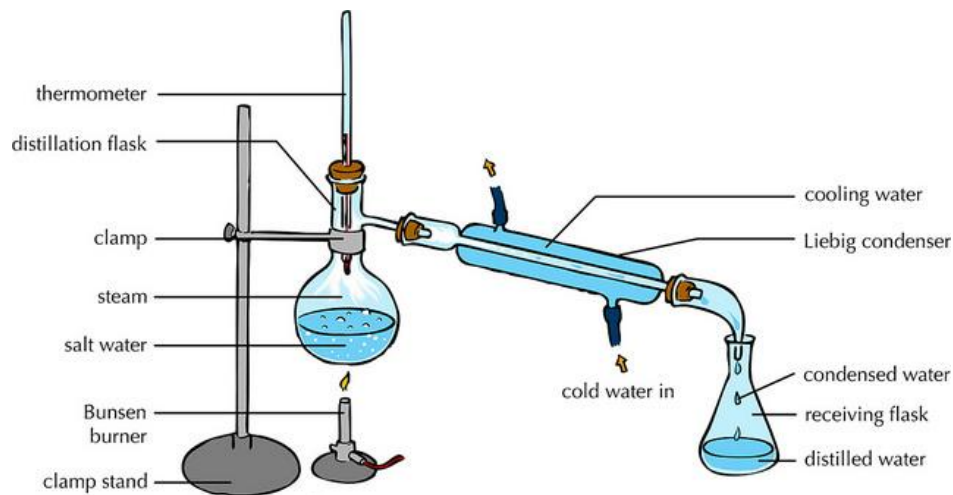
This technique separates a soluble substance from a solvent by evaporation



Example - crystallisation of sodium chloride from salt solution

Simple distillation

This technique separates a liquid from a mixture by evaporation followed by condensation

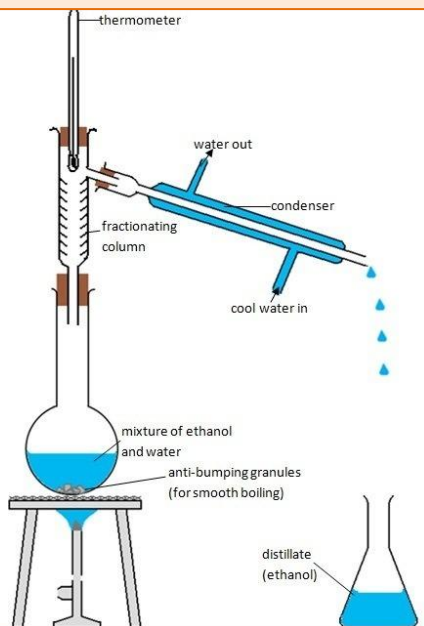


Example - obtaining water from sea water

Fractional distillation

This technique separates a mixture into a number of different parts, called **fractions**. Substances with **high** boiling points **condense at the bottom** and substances with **low** boiling points **condense at the top**.

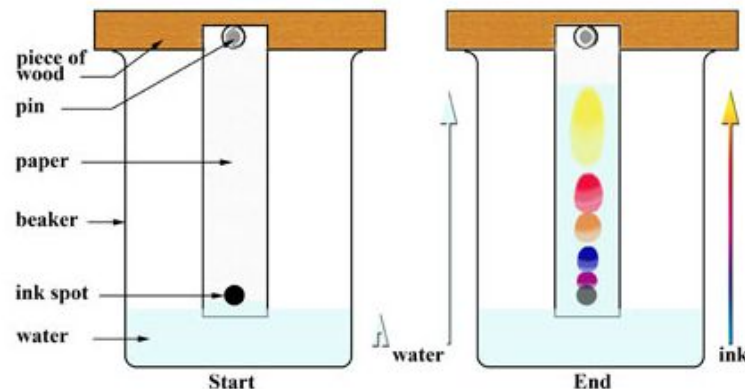
Fractional distillation works because the different substances in the mixture have different boiling points.



Example - obtaining ethanol from a mixture of ethanol and water

Chromatography

This technique separates small amounts of dissolved substances by running a solvent along absorbent paper



Example - separating the different colours in ink

QuestionIT

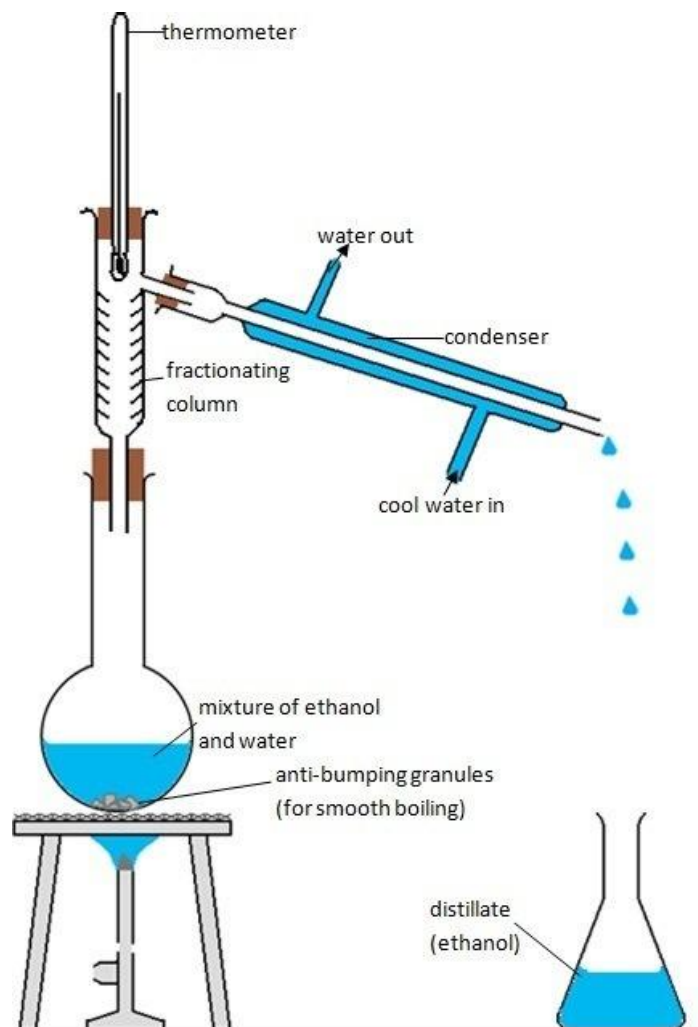
Atoms, elements,
compounds and mixtures
PART 2

- Mixtures

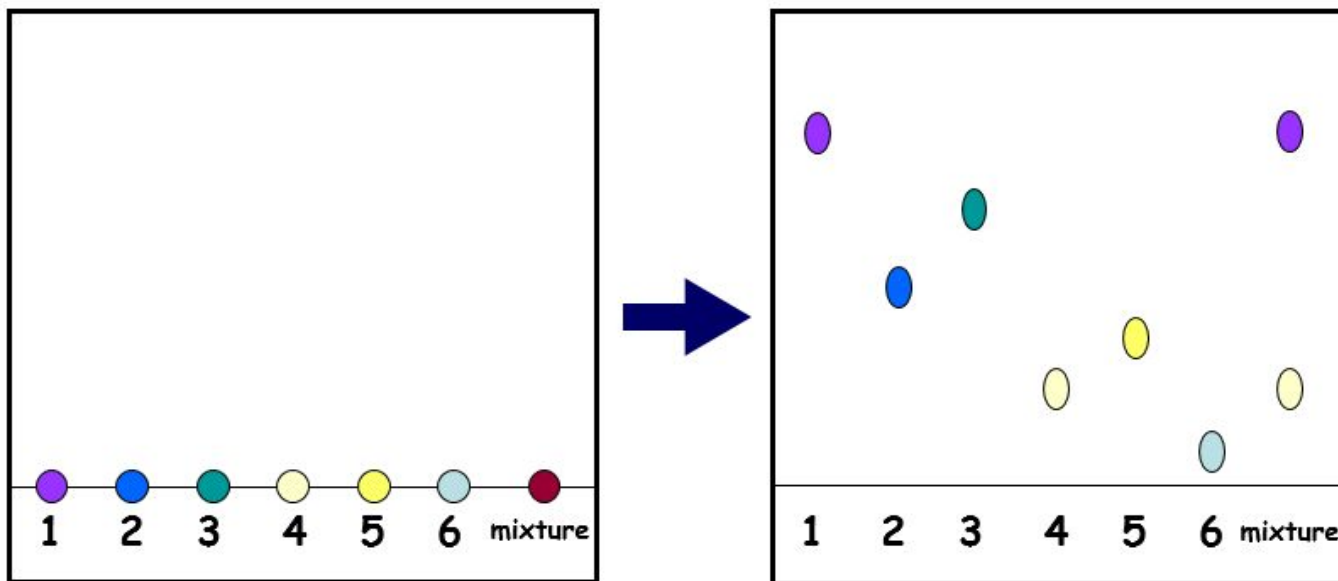


1. What is the difference between a compound and a mixture?
2. Describe the method used to collect the salt from a mixture of sand and salt.
3. What process is used to purify seawater to obtain usable water?

4. Describe how the following equipment is used to separate water and ethanol.



5. A mixture and six colours are tested using chromatography. The following chromatogram was produced.



- a) What can you conclude about the mixture?
- b) Why do the inks separate?

AnswerIT!

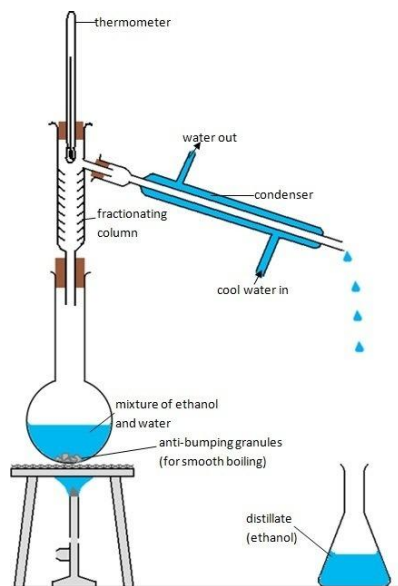
Atoms, elements,
compounds and
mixtures
PART 2

- Mixtures



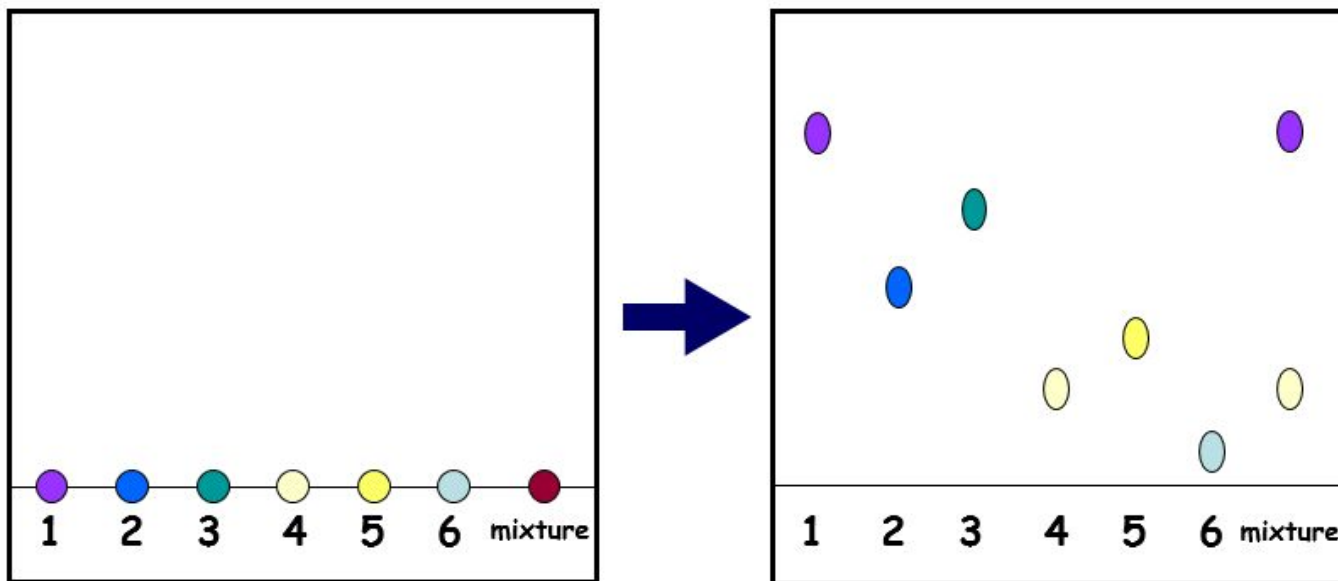
1. What is the difference between a compound and a mixture?
The substances in a mixture are not chemically bonded together, but they are in a compound.
2. Describe the method used to collect the salt from a mixture of sand and salt.
Add water and stir to dissolve the salt. Use filtration to remove the sand. Heat the water (gently) to allow the salt to crystallise (to avoid dryness).
3. What process is used to purify seawater to obtain usable water?
Simple distillation.

4. Describe how the following equipment is used to separate water and ethanol.



The ethanol and water mixture are heated. Ethanol has a boiling point of 78°C and will vaporise first. The gas passes over the fractionating column and into the condenser. Here it will condense, turning back into its liquid state.

5. A mixture and six colours are tested using chromatography. The following chromatogram was produced:



- a) What can you conclude about the mixture?

The mixture is made up from substance 1 and 4.

- b) Why do the inks separate to produce a chromatogram?

The separation depends on how soluble the chemical is in the solvent and how strongly the chemical is attracted to the paper.

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Atomic model PART 1

- Development of the atomic model

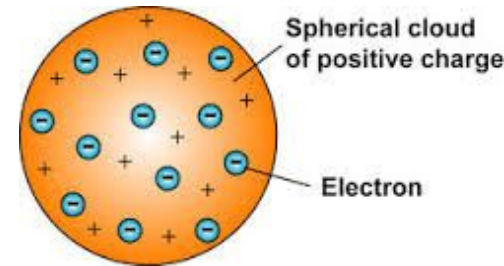


Early 1800s

Before the discovery of electrons, John Dalton's experiments led to the idea that atoms were **tiny spheres** that could **not be divided**.

End of 1800s

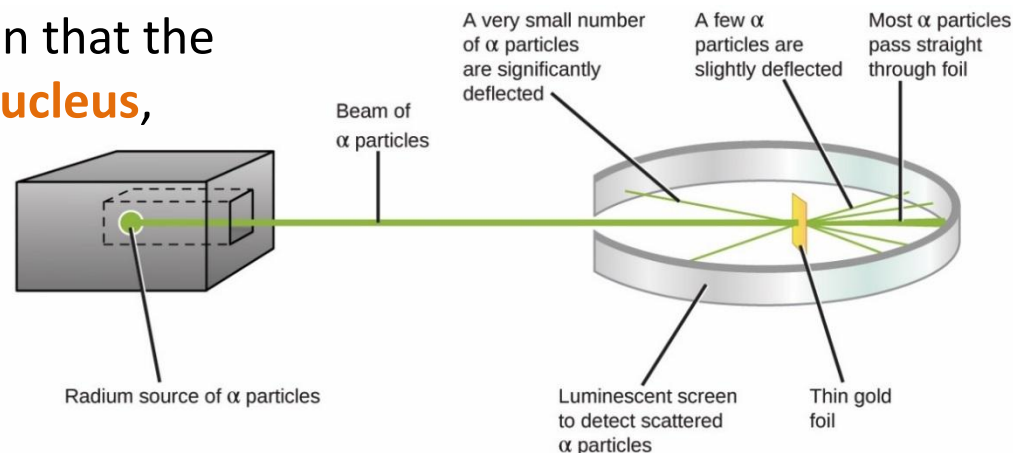
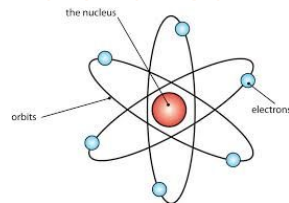
The **electron** was discovered by **JJ Thomson**. Scientists believed that atoms were **spheres of positive charge with negative charges spread** throughout - the '**plum-pudding**' model.



1908-1913

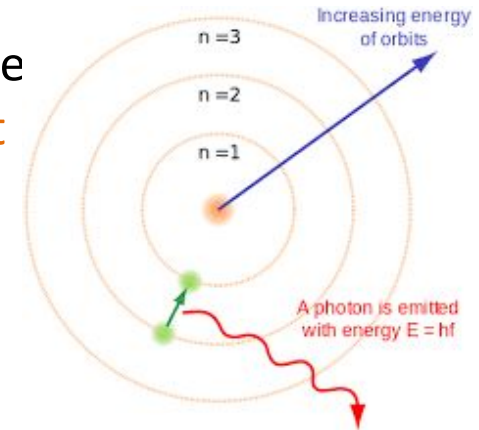
Ernst Rutherford designed an experiment carried out by **Geiger** and **Marsden**. They fired **alpha particles** at a piece of very thin **gold foil** (only a few atoms thick) which **scattered**, leading to the conclusion that the **mass of an atom was concentrated in a nucleus**, which was **charged**.

It proposed that **electrons orbited** around the nucleus.



1914

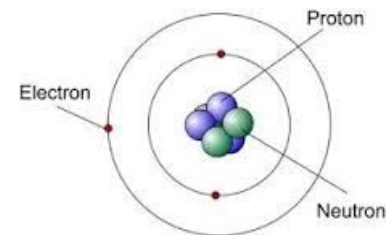
Niels Bohr noticed that the **light** given out when atoms were **heated** only had **specific** amounts of **energy** and he adapted the nuclear model by suggesting that **electrons orbit** the **nucleus** at **specific distances** in certain **fixed energy levels** (or shells). The energy must be given out when excited electrons **fall** from a **high** to **low** energy level.



Later experiments led to the idea that the **positive charge** of the **nucleus** could be **subdivided** into a **whole number** of **smaller particles**, each particle having the **same** amount of **positive charge**. The name **proton** was given to these particles.

1932

James Chadwick bombarded **beryllium** atoms with **alpha** particles. An unknown radiation was produced. Chadwick interpreted this radiation as being composed of particles with a **neutral electrical charge** and the approximate **mass of a proton**. This particle became known as the **neutron**.



- John Dalton – **tiny spheres that could not be divided.**

- JJ Thomson – **electron** discovered. **Plum pudding model** - spheres of positive charge with negative charges spread evenly though.

- E. Rutherford, Geiger and Marsden - **alpha particle scattering experiment**. Nuclear model - mass of atom concentrated in a **charged nucleus**, with **orbiting electrons**.

- **Niels Bohr** – **electrons orbit nucleus at specific distances in fixed energy levels** (shells). Energy given out when electrons change level.

- Positive charge of nucleus could be subdivided into particles of positive charge – **protons**.

- **James Chadwick** – provided evidence for the existence of **neutrons** within the nucleus.

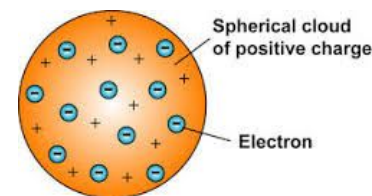
QuestionIT

Atomic model PART 1

- Development of the atomic model



1. What was the earliest model of the atom?
2. Which subatomic particle did JJ. Thomson discover?
3. Which early atomic model does the following diagram show?
4. Name Rutherford's experiment.
5. State two ways in which Rutherford's experiment changed Thomson's model of the atom.
6. How did Bohr adapt the nuclear model?
7. Explain why Bohr revised Rutherford's model of the atom.



AnswerIT!

Atomic model PART 1

- Development of the atomic model



1. What was the earliest model of the atom?

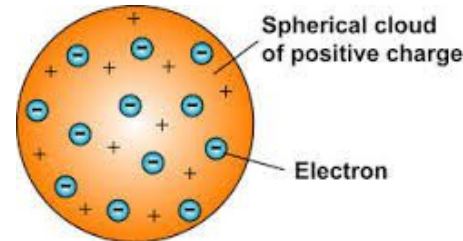
Tiny spheres that could not be divided.

2. Which subatomic particle did JJ. Thomson discover?

The electron

3. Which early atomic model does the following diagram show?

The plum pudding model.



4. Name Rutherford's experiment.

Alpha particle scattering.

5. State two ways in which Rutherford's experiment changed Thomson's model of the atom.

He said the positive charge was concentrated into very small volume at centre of atom (nucleus) and the electrons orbit nucleus.

6. Explain why Bohr revised Rutherford's model of the atom.
Suggesting electrons orbit the nucleus at different distances.

7. Explain why Bohr revised Rutherford's model of the atom.
Energy emitted from electron transitions can only have certain fixed energies, so he refined the 'orbiting electrons' in Rutherford's nuclear model to 'orbiting electrons in energy levels (or shells) at fixed distances from the nucleus'.

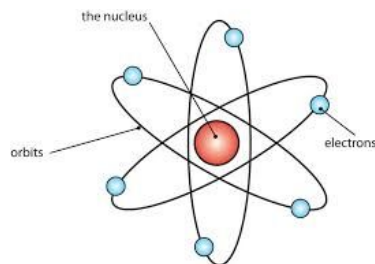
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Atomic model PART 2

- Relative electrical charges of subatomic particles.
- Size and mass of atoms
- Atomic structure



Subatomic particles



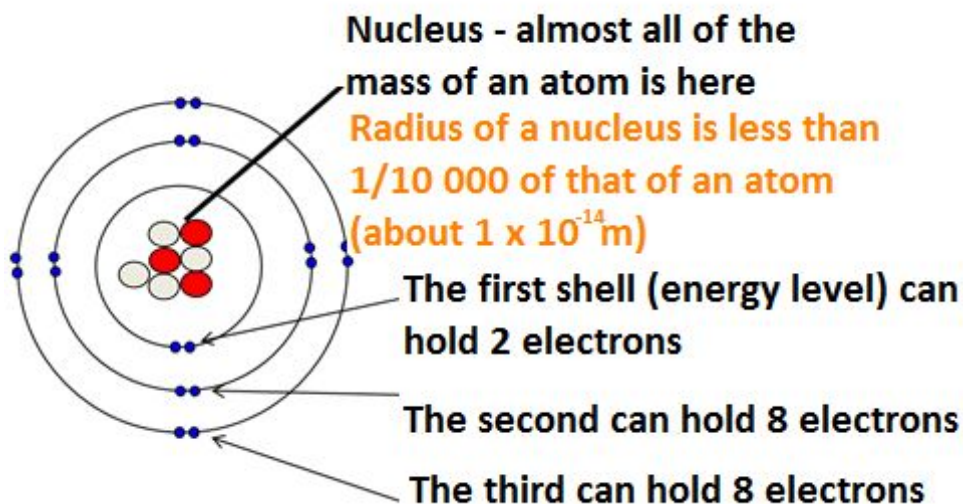
	Mass	Charge	Location
Proton	1	+	nucleus
Neutron	1	0	nucleus
Electron	Very small	-	shells

- An atom contains **equal** numbers of **protons** and **electrons**.
- Atoms have **no** overall electrical charge because the number of positive protons equals the number of negative electrons.

Number of protons = atomic number.

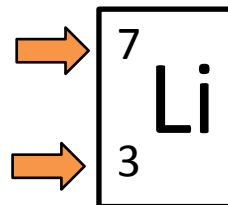
- All atoms of an element have the **same** number of **protons**.
- Atoms of **different** elements have **different numbers** of **protons**.

Atoms are very small, having a radius of about **0.1nm** ($1 \times 10^{-10}\text{m}$).
Protons and Neutrons are found in the **nucleus**. Electrons orbit the nucleus in **shells**.



Mass number = Number of **protons** and **neutrons**

Atomic number = Number of protons



Protons = 3

Electrons = 3

Neutrons = 4

To calculate the number of neutrons = **Mass Number** – **Atomic Number**

Atoms sometimes **lose or gain** electrons (e.g. when a metal reacts with a non-metal). When they do this they become a charged atom or an **ion**.

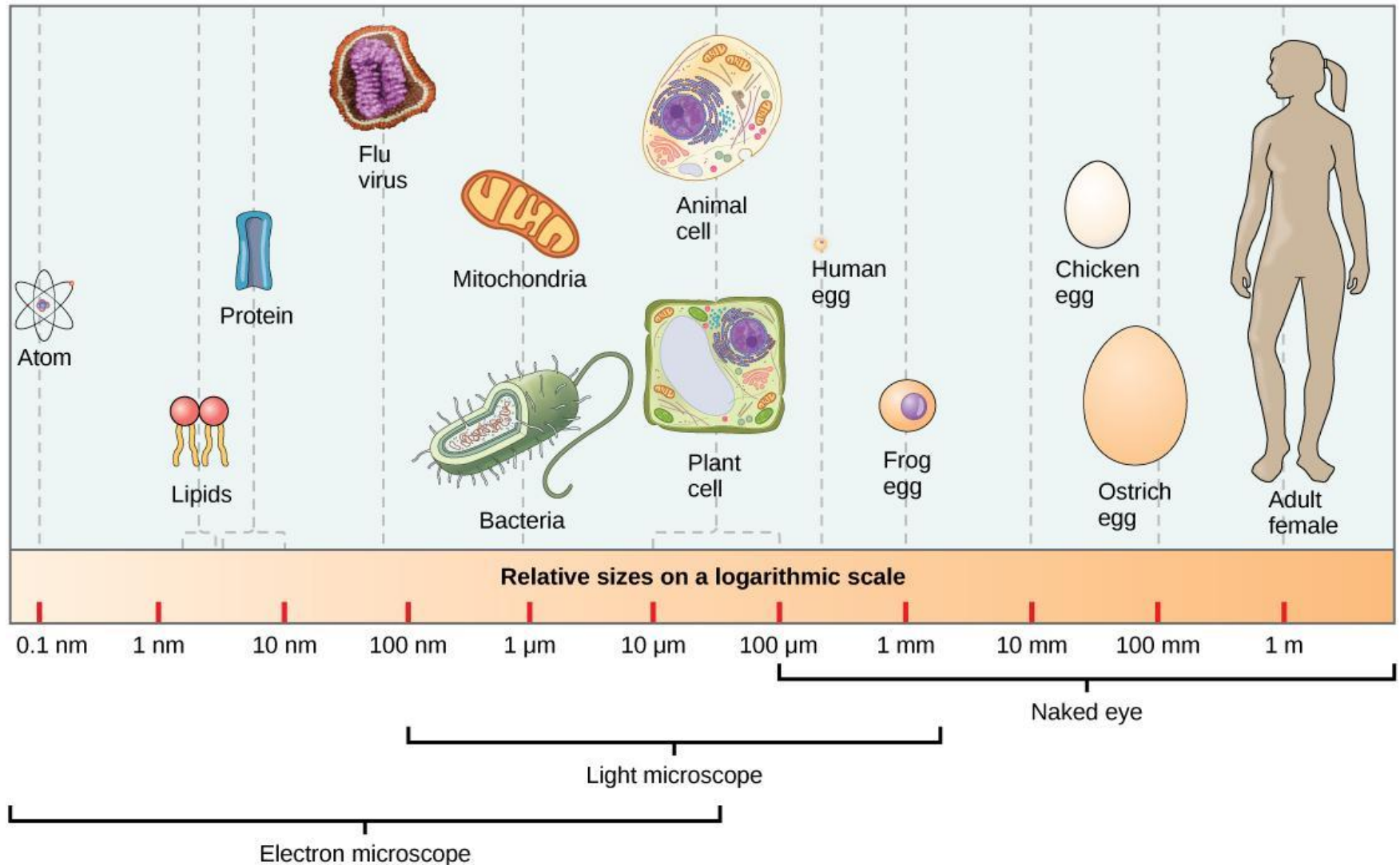
If an atom loses one or more electrons, it gains a **positive** charge because it has less electrons than protons.

e.g. If sodium atom loses one electron, it forms a **Na⁺ ion**.
It has 11 protons, 12 neutrons and 10 electrons.

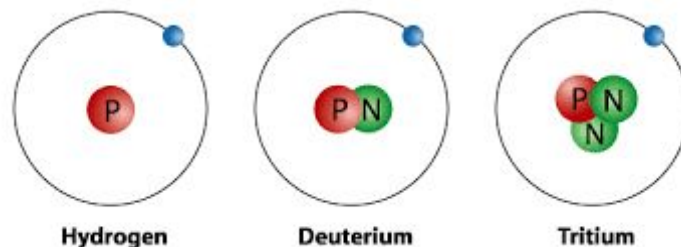
If an atom gains one or more electrons, it gains a **negative** charge because it has more electrons than protons.

e.g. If an oxygen atom gains two electrons, it forms a **O²⁻ ion**.
It has 8 protons, 8 neutrons and 10 electrons.

Atoms are tiny - the radius of a typical atom is **one tenth** of a billionth of a meter.



Atoms of the **same element** can have **different numbers** of **neutrons** – an **isotope**.



Isotopes of Hydrogen

The **relative atomic mass** of an element is an **average** value that takes account of the **abundance** of the isotopes of the element. Samples of **different** isotopes of an element have different **physical properties** (e.g. different density), however, they always have the **same chemical properties**. It is calculated by working out the relative **abundance** of each isotope.

The relative atomic mass is therefore calculated using the equation:

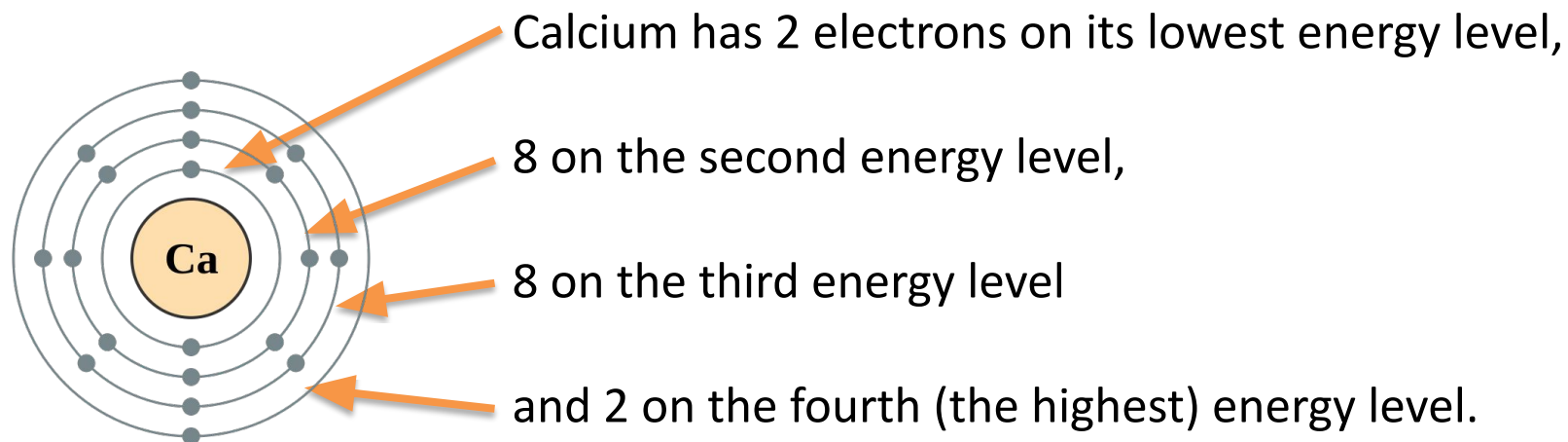
$$(\% \text{ of isotope 1} \times \text{mass of isotope 1}) + (\% \text{ of isotope 2} \times \text{mass of isotope 2}) \div 100$$

Chlorine's relative atomic mass of 35.5 is an average of the masses of the different isotopes of chlorine.

For example, in any sample of Chlorine 25% will be ^{37}Cl and 75% ^{35}Cl .

$$(25 \times 37) + (75 \times 35) \div 100 = 35.5$$

The electrons in an atom occupy the lowest available **energy levels** (innermost available shells). The electronic structure of an atom can be represented by numbers or by a diagram e.g.



So the electron configuration for Calcium = 2,8,8,2

QuestionIT

Atomic model PART 2

- Relative electrical charges of subatomic particles.
- Size and mass of atoms
- Atomic structure



1. Name three subatomic particles and their charges.
2. Complete the sentence 'All atoms of one type of element have the same number of...'
3. What does the atomic number tell us about an atom?
4. What does the mass number tell us about an atom?
5. How is an isotope different to an atom?
6. What is the electron configuration for sodium?

7. Why is the overall charge of an atom zero?
8. Draw the electron configuration for a chlorine atom.
9. Write the electron configuration for potassium.
10. How many electrons does potassium have on its highest energy level?

11. Calculate how many protons, electrons and neutrons there are in
- A silver atom with atomic number 47 and mass number 108
 -

12. The atom ${}_{19}^{39}\text{K}^{+}$ radius of a bromine atom is $9 \times 10^{-11}\text{m}$
- Give its atomic radius in nanometres
 - Calculate the radius of the nucleus (in nm), given that it will be about 1/10,000 the radius of the atom. Give your answer in standard form.

AnswerIT!

Atomic model PART 2

- Relative electrical charges of subatomic particles.
- Size and mass of atoms
- Atomic structure

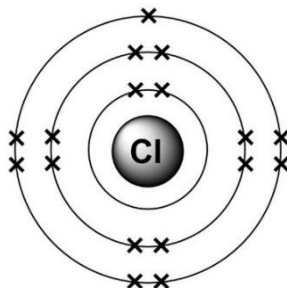


1. Name three subatomic particles and their charges.
Proton – positive; neutron – no charge; electron – negative.
2. Complete the sentence 'All atoms of one type of element have the same number of...'
Protons.
3. What does the atomic number tell us about an atom?
Number of protons.
4. What does the mass number tell us about an atom?
Number of protons + number of neutrons.
5. How is an isotope different to an atom?
Different number of neutrons.
6. What is the electron configuration for sodium?
2,8,1

7. Why is the overall charge of an atom zero?

Number of protons (positive) = number electrons (negative).

8. Draw the electron configuration for a chlorine atom.



9. Write the electron configuration for potassium.

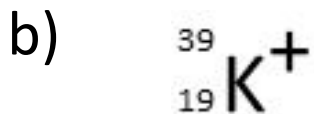
2,8,8,1

10. How many electrons does potassium have on its highest energy level?

1

11. Calculate how many protons, electrons and neutrons there are in:
a) A silver atom with atomic number 47 and mass number 108

Protons = 47 Electrons = 47 Neutrons = 61



Protons = 19 Electrons = 18 Neutrons = 20

12. The atomic radius of a bromine atom is $9 \times 10^{-11}\text{m}$

- a) Give its atomic radius in nanometres.

0.09 nm

- b) Calculate the radius of the nucleus (in nm), given that it will be about 1/10,000 the radius of the atom. Give your answer in standard form.

$9 \times 10^{-6}\text{nm}$

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Periodic table PART 1

- The periodic table
- Development of the periodic table
- Metals and non-metals



The elements are arranged in order of increasing **atomic number**.

1	2											3	4	5	6	7	0	
H																		He
Li	Be											B	C	N	O	F	Ne	
Na	Mg											Al	Si	P	S	Cl	Ar	
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	?	?	?							

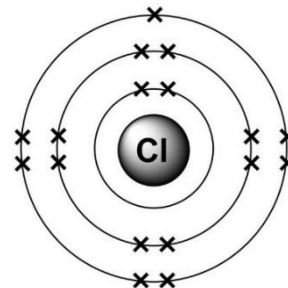
Elements with similar properties are in **columns**, known as **groups**.

Elements in the same group have the same number of electrons in their outer shell.

The **rows** in the table are called **periods**

It is called a periodic table because similar properties occur at regular intervals

Group = electrons in outer shell
Period = number of shells



Group = 7
Period = 3

1808

John Dalton published a table of elements that were arranged in order of their atomic weights, which had been measured in various chemical reactions

ELEMENTS					
	Hydrogen.	1 ^{wt}		Strontian	46 ^{wt}
	Azote	5		Barytes	68
	Carbon	5 ¹		Iron	56
	Oxygen	7		Zinc	56
	Phosphorus	9		Copper	56
	Sulphur	13		Lead	90
	Magnesia	20		Silver	190
	Lime	24		Gold	190
	Soda	28		Platina	190
	Potash	42		Mercury	167

1864

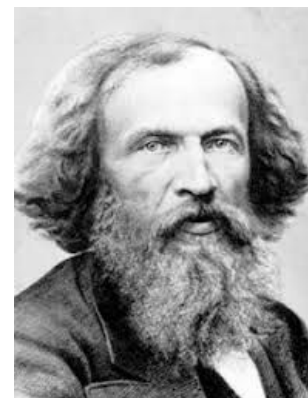
John Newlands published the law of octaves. However the **table was incomplete and elements were placed in inappropriate groups**

Newlands' Arranged Elements in Octaves:

H	F	Cl	Co/Ni	Br	Pd	I	Pt/Ir
Li	Na	K	Cu	Rb	Ag	Cs	Tl
G	Mg	Ca	Zn	Sr	Cd	Ba/V	Pb
Bo	Al	Cr	Y	Ce/La	U	Ta	Th
C	Si	Ti	In	Zn	Sn	W	Hg
N	P	Mn	As	Di/Mo	Sb	Nb	Bi
O	S	Fe	Se	Ro/Ru	Te	Au	Os

1869

Dmitri Mendeleev overcame Dalton's problem by **leaving gaps for the elements that he thought had not been discovered** and **in some places changed the order based on atomic weight** (e.g. Argon and Potassium). Elements with properties predicted by Mendeleev were eventually discovered.



Early 20th Century - Scientists began to find out more about the atom and knowledge of **isotopes** explained why the order was not always correct.

The elements can be divided into **metals** and **non-metals**.

[illegible]

Elements that do not form positive ions are non-metals

Elements that tend to form positive ions are metals

Non metals – found towards the right and towards the top of the periodic table

**Most elements are metals –
found towards the left and
towards the bottom of the
periodic table**

Metals	Non-metals
Shiny	Dull
Mostly solid	Low density
Dense and strong	Weak
Malleable	Brittle
Good heat and electrical conductors	Poor heat and electrical conductors

QuestionIT

Periodic table PART 1

- The periodic table
- Development of the periodic table
- Metals and non-metals



1. How are elements arranged in the periodic table?
2. What are the columns of the periodic table called?
3. What are the rows of a periodic table called?
4. What does the column an element is in tell you about the atoms?
5. What does the row an element is in tell you about the atoms?
6. Which side of the periodic table are the non-metals found?

7. How were elements classified before the discovery of subatomic particles?
8. What was the problem with early periodic tables?
9. How did Mendeleev overcome these problems?
10. Why is the order based on atomic masses not always correct?
11. What do we call elements that tend to form positive ions?

AnswerIT!

Periodic table PART 1

- The periodic table
- Development of the periodic table
- Metals and non-metals



1. How are elements arranged in the modern periodic table?

Increasing atomic number.

2. What are the columns of the periodic table called?

Groups.

3. What are the rows of a periodic table called?

Periods.

4. What does the column an element is in tell you about the atoms?

Number of electrons in outer shell.

5. What does the row an element is in tell you about the atoms?

Number of electron shells.

6. Which side of the periodic table are the non-metals found?

Right and top.

7. How were elements classified before the discovery of subatomic particles?

Arranging in order of their atomic weights.

8. What was the problem with early periodic tables?

Incomplete; some elements placed in inappropriate groups.

9. How did Mendeleev overcome these problems?

Left gaps for elements he thought had not yet been discovered; changed the order based on atomic weights for some elements if they did not fit the pattern.

10. Why is the order based on atomic masses not always correct?

The presence of isotopes.

11. What do we call elements that tend to form positive ions?

Metals.

LearnIT! KnowIT!

Periodic table PART 2

- Group 0, Group 1
and Group 7



1	2		3	4	5	6	7	0
H								He
Li	Be		B	C	N	O	F	Ne
Na	Mg		Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Cobalt
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh
Cs	Ba	Lanthanum	Hf	Ta	W	Re	Os	Ir
Fr	Ra	Actinium	Rf	Db	Sg	Bh	Hs	Mt

THE NOBLE GASES

Elements in **Group 0** of the periodic table are called the **noble gases**. They are unreactive because their atoms have **stable** arrangements of **electrons**. The atoms have **eight** electrons in their outermost shell, apart from helium, which has just **two**, but still has a complete outer shell.

The stable electronic structure explains why they exist as **single atoms**, they have no tendency to react to form **molecules**.

The **boiling points** of the noble gases get **higher** going **down** the group. For example helium boils at -269°C and radon boils at -62°C .

THE ALKALI METALS

[illegible]

The alkali metals are very reactive. They need to be stored under **oil** to prevent them reacting with **oxygen** and **water** vapour in the **air**. The alkali metals have **low densities**. The metals are very **soft** and can be **cut** with a knife. They also have **low melting** and **boiling points**.

The properties are due to all the atoms having just **one** electron in their outermost shell. They only need to **lose** one electron to get the stable electronic structure of a noble gas.

Li
Na
K
Rb
Cs

Reactivity Increases

The atoms get **larger** as you go **down**, so the **single** electron in the outermost shell (highest energy level) is attracted **less** strongly to the positive nucleus. The **electrostatic attraction** with the nucleus gets **weaker** because the **distance** between the outer electron and the nucleus **increases**. Also the outer electron experiences a **shielding effect** from the inner electrons, **reducing** the attraction between the oppositely charged outer electron and the nucleus.

The alkali metals have a **silvery, shiny** surface when they are first cut. However, this goes **dull** very quickly as the metals reacts with the oxygen in the air.

e.g. **sodium + oxygen** \rightarrow **sodium oxide**



Lithium, sodium and potassium all react **vigorously** with water. When you add them to water, the metal **floats, moves** around and **fizzes**.

e.g. **potassium + water** \rightarrow **potassium hydroxide + hydrogen**



Potassium ignites with a lilac flame



They also react vigorously with non metals, such as group seven. They react to make **ionic compounds**. These are generally soluble in water, **giving colourless solutions**.

sodium + chlorine \rightarrow **sodium chloride**



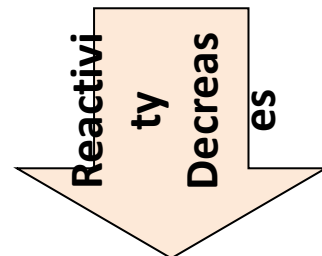
THE HALOGENS

When Group 7 elements react, the atoms **gain** an electron in their outermost shell. Going down the group, the outermost shell's electrons get **further away** from the attractive force of the nucleus, so it is **harder** to **attract** and **gain** an extra electron. The outer shell will also be **shielded** by more inner shells of electrons, again **reducing** the **electrostatic attraction** of the nucleus for an incoming electron.

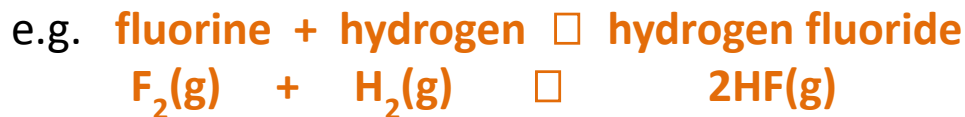
The halogens are a group of **toxic** non-metals that have **coloured** vapours. They have **low melting** and **boiling** points, which **increase** down the group. They are **poor conductors** of **heat** and **electricity**.

As elements, the halogens exist as **molecules** made up of pairs of atoms. These are called **diatomic** molecules **F₂, Cl₂, Br₂, I₂ and At₂**. The halogens have **seven** electrons in their outermost shell and need to **gain** one electron to achieve the stable electronic structure of a noble gas. When they react with non metals, they are joined together by a **covalent** bond.

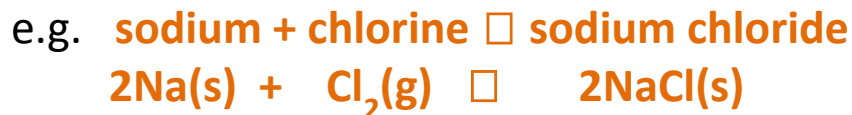
F
Cl
Br
I
At



The halogens react with **hydrogen**. The reactions with hydrogen become **less** reactive as you go **down** the group.



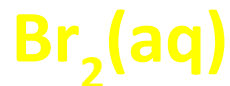
The halogens also react with **metals**. The halogen atoms **gain** a single electron to give them a stable arrangement of electrons. They form **ionic compound**.



A more reactive halogen will also **displace** a less reactive halogen from solutions of its **salts**.



The **colour** of the solution after mixing depends on the **less** reactive pair of halogens.



QuestionIT!

Periodic table PART 2

- Group 0, Group 1
and Group 7



1. What are the elements in Group 0 called?
2. What are the elements in Group 1 called?
3. What are the elements in Group 7 called?
4. What happens to the boiling point of elements in Group 0 as you go down the group?
5. Why are the elements in Group 0 so unreactive?
6. Why do all elements in Group 1 react in a similar way to each other?

7. What happens to the reactivity of the elements as you go down Group 1?
8. Write a word equation for the reaction between sodium and oxygen.
9. Why do all the elements in Group 7 react in a similar way to each other?
10. Halogens are diatomic. What does the word 'diatomic' mean?
11. What happens to the reactivity as you go down Group 7?

12. What happens to the melting point and boiling point as you go down Group 7?
13. Write a word equation for the reaction between lithium and chlorine.
14. Write a balanced chemical equation for the reaction between lithium and chlorine gas.

AnswerIT!

Periodic table PART 2

- Group 0, Group 1
and Group 7



1. What are the elements in Group 0 called?
Noble gases.
2. What are the elements in Group 1 called?
Alkali metals.
3. What are the elements in Group 7 called?
Halogens.
4. What happens to the boiling point of elements in Group 0 as you go down the group?
Increase.
5. Why are the elements in Group 0 so unreactive?
Full outer shell of electrons.
6. Why do all elements in Group 1 react in a similar way to each other?
1 electron in outer shell.

7. What happens to the reactivity of the elements as you go down Group 1?

Increases.

8. Write a word equation for the reaction between sodium and oxygen.

sodium + oxygen \rightarrow sodium oxide

9. Why do all the elements in Group 7 react in a similar way to each other?

7 electrons in outer shell.

10. Halogens are diatomic. What does the word 'diatomic' mean?

There molecules contain 2 atoms.

11. What happens to the reactivity as you go down Group 7?

Decreases.

12. What happens to the melting point and boiling point as you go down Group 7?

Increases.

13. Write a word equation for the reaction between lithium and chlorine.

lithium + chlorine \rightarrow lithium chloride

14. Write a balanced chemical equation for the reaction between lithium and chlorine gas.

$2\text{Li(s)} + \text{Cl}_2\text{(g)} \rightarrow 2\text{LiCl(s)}$

LearnIT! KnowIT!

Periodic table PART 3

- Transition Metals
(chemistry only)



Chemistry only

THE TRANSITION METALS

[illegible]

The general trend is the reactivity **decreases across** the **period**, but there are exceptions, for example Zinc is very reactive.

Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
----	----	---	----	----	----	----	----	----	----

Reactivity decreases

The transition metals are located **between group 2 and group 3.**

The transition metals have:

High melting points

High boiling points

High densities

They are:

Shiny when polished

Malleable – can be hammered into a shape

Strong – don't break easily when a force is applied

Sonorous – makes a ringing sound when hit

Ductile – can be stretched into wires
Conductors of **electricity** and **heat**

	Alkali metals	Transition metals
Melting points	Low	High (except mercury, which is liquid at room temperature)
Reactivity	High (react vigorously with water or oxygen)	Low (do not react so vigorously with water or oxygen)
Strength	Soft or liquid so cannot withstand force	Strong and hard
Density	Low	High
Compounds formed	White or colourless	Coloured

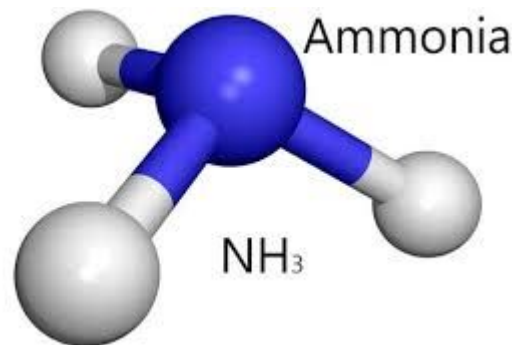
The transition metals have many different **uses** due to their **properties**. **Copper** has **properties** that make it useful for **electrical wiring** and **plumbing**. Not very reactive, **excellent conductor of electricity**, easily **bent** into shape for **water pipes** in plumbing.

They can also be useful as **catalysts**. A catalyst is a substance that **speeds up** a **chemical reaction** without being used up. Catalysts are hugely valuable in industry where they can **save time** and **energy**.

Nickel is the catalyst used in the hydrogenation of oil to produce **margarine**



Iron is the catalyst used in the Haber process to produce **ammonia**



QuestionIT

Periodic table PART 3

- Transition Metals
(chemistry only)



1. Where are transition metals found on the periodic table?
2. How do the melting points of transition metals compare to Group 1 metals?
3. How do the densities of transition metals compare to Group 1 metals?
4. How does the strength of transition metals compare to Group 1 metals?

5. Describe the differences between the reactions of the alkali metals and the reactions of transition metals.
6. State two typical properties of transition metals.
7. State one use of transition metals.
8. Explain why copper is used for plumbing.

AnswerIT!

Periodic table PART 3

- Transition Metals
(chemistry only)



1. Where are transition metals found on the periodic table?
Between Group 2 and Group 3.
2. How do the melting points of transition metals compare to Group 1 metals?
Transition metals have a higher melting point than alkali metals.
3. How do the densities of transition metals compare to Group 1 metals?
Transition metals are more dense than alkali metals.
4. How does the strength of transition metals compare to Group 1 metals?
Transition metals are stronger than alkali metals.

5. Describe the differences between the reactions of the alkali metals and the reactions of transition metals.

Alkali metals react vigorously with water and with oxygen from the air. The transition elements react slowly with these reagents if at all.

6. State two typical properties of transition metals.

Have ions with different charges; form coloured compounds.

7. State one use of transition metals.

Transition metals can be used as catalysts.

8. Explain why copper is used for plumbing.

It has a high melting point, it conducts heat and is strong and malleable. It also does not react with water.