GCSE

Introduction to the Periodic Table

Group Number	1	2				G	R	O	U	P			3	4	5	6	7	8 or 0
Group Name	Alkali Metals	parts of the PERIODIC TABLE atomic numbers 1-38, 49-54															The Halogens	Noble Gases
Period 1							H 1											Helium 2, He
Period 2	Lithium 3, Li	Be 4											B 5	C 6	N 7	O 8	Fluorine 9, F	Neon 10, Ne
Period 3	Sodium 11, Na	Mg 12											Al 13	Si 14	P 15	S 16	Chlorine 17, Cl	Argon 18, Ar
Period 4	Potassium 19, K	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Bromine 35, Br	Krypton 36, Kr
Period 5	Rubidium 37, Rb	Sr 38	T	R	A M	N E	S T	I A	T L	I S	O	N	In 49	Sn 50	Sb 51	Te 52	Iodine 53, I	Xenon 54, Xe

- The elements are laid out in order of **Atomic (proton) Number** (at. no.). **Originally they were laid out in order of 'relative atomic mass'** (the old term was 'atomic weight').
- **Many of the similarities and differences in the properties of elements can be explained by the electronic structure of the atoms (electron configuration, arrangement in shells or energy levels)**.
- The idea of the Periodic Table is to arrange the elements in a way that enables chemist's to understand patterns in the properties of the elements.
- **The main structural features of the periodic table are** ...
  - to produce **columns of similar elements called Groups.**
    - They are usually similar chemically and physically BUT there are often important trends in physical properties and chemical reactivity.
  - The resulting **complete horizontal rows are called Periods** and usually consist of a range of elements of different character from metals on the left to non-metals on the right.
    - BUT within a period you can get a series of like elements eg the **1st Transition Series of Metals** (Sc to Zn) in Period 4.
  - The ideas of **Group and Period are totally connected with electron structure** (see below)

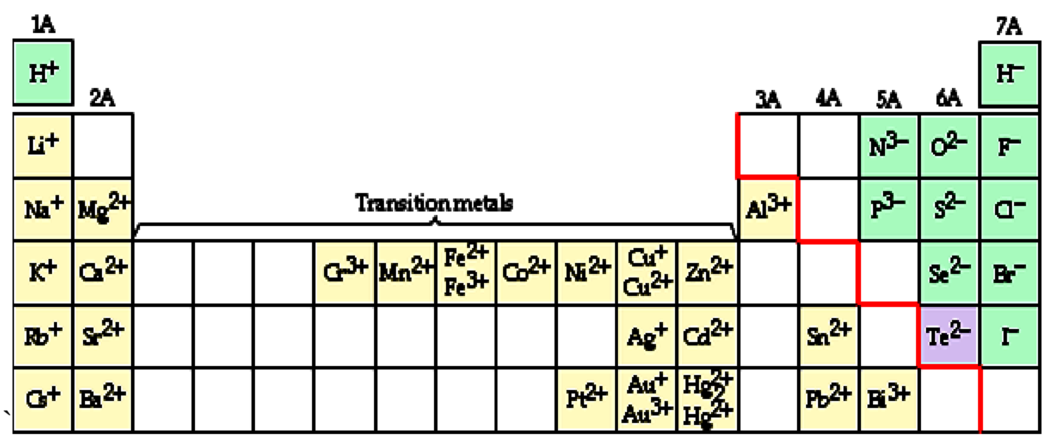
Group number and Name	Group 1 The Alkali Metals	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7 The Halogens	Group 0 Noble Gases
Period 1	$_1$H hydrogen doesn't really fit in any group The electron arrangements of the first twenty elements							$_2$He 2
Period 2	$_3$Li 2.1	$_4$Be 2.2	$_5$B 2.3	$_6$C 2.4	$_7$N 2.5	$_8$O 2.6	$_9$F 2.7	$_{10}$Ne 2.8
Period 3	$_{11}$Na 2.8.1	$_{12}$Mg 2.8.2	$_{13}$Al 2.8.3	$_{14}$Si 2.8.4	$_{15}$P 2.8.5	$_{16}$S 2.8.6	$_{17}$Cl 2.8.7	$_{18}$Ar 2.8.8
Period 4	$_{19}$K 2.8.8.1	$_{20}$Ca 2.8.8.2						

All substances are made up of one or more of the different types of atoms we call elements.

- **Hydrogen, 1, H, the simplest element atom,** does not readily fit into any group.
- A **Group** is a vertical column of like elements eg **Group 1 The Alkali Metals** (Li, Na, K etc.), **Group 2 The Alkaline Earth Metals, (Ca, Mg etc.)** **Group 7 The Halogens** (F, Cl, Br, I etc.) and **Group 0 (8) The Noble Gases** (He, Ne, Ar etc.).
- Apart from hydrogen (doesn't really fit in any group), and helium (*****), the **group number equals the number of electrons in the outer shell** (eg chlorine's electron arrangement is 2.8.7, the second element down in Group 7 on period 3). So , after helium, elements in the same group have the same outer electron structure.
- **The elements in a group tend to have similar physical and chemical properties because of their similar outer shell electron structure**.
- (***** although helium can't have 8 outer electrons like the rest of Group 0, its outer shell of 2 electrons is complete, just like neon and argon etc.)
- A **Period is a horizontal row of elements** with a variety of properties, changing from very metallic elements on the left to non-metallic elements on the right. A period starts when the next electron goes into the next available main energy level or shell (Group 1 alkali Metals). The period ends when the main energy level is full (Group 0 or 8 Noble Gases).
- **All the elements on the same period use the same number of principal electron shells**, and this equals the period number (eg sodium's electron arrangement 2.8.1, the first element in Period 3).
- The first element in a period is when the next electron goes into the next available electron shell or energy level (ie 1 electron in the outer shell, after H it is the Group 1 Alkali Metals like sodium 2.8.1).
- **The last element in a period is when the outer shell is full** (The Group 0 Noble Gases eg argon 2.8.8). The next electron for the next element goes into the next highest level (shell) available, and so starts the next period.
- So in terms of electrons ....
    - Period 1 is elements 1-2, H (1) to He (2)
    - Period 2 is elements 3-8, Li (2.1) to Ne (2.8)
    - Period 3 is elements 11-18, Na (2.8.1) to Ar (2.8.8)
    - Period 4 is elements 19-36, starts with K (2.8.8.1) and Ca (2.8.8.2) and finishes with the Noble Gas Kr (2.8.18.8).
    - Note that the number of shells containing electrons is equal to the period number.
- The similarities (eg same Group) or differences (eg across a period) of the properties of the elements can be explained by the electronic structure of the atoms.
- From Period 4 onwards the length of a period significantly increases because it includes horizontal series of similar metals with their own characteristic physical and chemical properties eg **The 1st Transition Metals Series**.
- More than **three-quarters** of the **109 known elements** are **metals** (elements **naturally occur up to uranium 92**, **93-109 are 'man-made' elements** from the experiments of nuclear physicists.
    - This work will continues as heavier and heavier elements are likely to be made in nuclear reactions. They will be all metals and radioactive. BUT one theory suggests that 'super-heavy' elements of about atomic number 150? may be in a nuclear stability region and would prove most interesting to study. Chemists are trying to predict their properties now!, so it may have started with Mendeleev but it ain't finished yet!
- **Only about 19 are definitely are non-metal** but **about 7 more are semi-metals of mixed physical and chemical character**.
- **The metals in the periodic table are mainly found in the left hand columns (Groups 1 and 2) and in the central blocks of the transition elements**.
- There is a 'rough' diagonal division between the two principal types of element zig-zagging from B-Al in group 3 to Te-Po in Group 6.
- The elements in this 'band' are sometimes referred to as '**semi-metals**' or '**metalloids**' because of **their 'mixture' of metallic and non-metallic physical or chemical character** eg the semi-conductor silicon in group 4.
- **There tends to be gradual changes in physical and chemical properties down a group** eg
    - Down Group 1 (Alkali Metals) and Group 2 the metals get more reactive.
    - Down Group 7 (Halogens) the non-metals get less reactive, their colour gets darker, their melting/boiling points increase.
- **There tends to be major changes in physical and chemical properties across a period** eg

- Period 2 starts with a solid low melting reactive metal lithium, in the middle there are the high melting and rather unreactive non-metals boron and carbon, next to the end is the very highly reactive non-metal gas fluorine, and the period finishes with the very unreactive gas neon.
- From left to right across a period the bonding in chlorides or oxides changes from ionic (with metals eg Na^+Cl^-, $(Na^+)_2O^{2-}$) to covalent (with non-metals eg ClF, SO_2).
- From left to right across a period the oxides change from alkaline/basic (with metals eg Na_2O) to acidic (with non-metals eg SO_2)
- **Note on electron arrangements**:
  - Except for boron, **most non-metals** have at least four electrons in the out shell.
  - Except for the noble gases, the more electrons in the outer shell the more non-metallic and the more reactive the element. The **most reactive non-metals only need to share/gain one or two electrons**.
  - The **most reactive metals** only have 1 or 2 electrons in the outer shell which tend to be easily lost to form the metal ion in reaction.
  - The most reactive metals have a low number of outer valency shell electrons (<= 3).
  - The very reactive non-metals have 5 to 7 outer valency shell electrons.
  - Elements in the 'middle' of the Periodic Table eg Group 4 with 4 outer electrons, show mixed chemical character and are not very reactive elements.
  - The Noble Gas elements have full, very stable, outer valency shells.

Periodic Table of Selected Ions



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Typical Properties of Metallic Elements

Physical properties of metals

- **high melting points and boiling points** so all solid bar one (exceptions like mercury the only liquid metal at room temperature and the Alkali Metals have untypical low melting points)
- **good conductors of heat and electricity**
- **high density** (exceptions like the Alkali Metals have untypical low densities, the first three Li, Na and K float on water before the 'fizzing'!)
- **appearance - shiny** (usually silvery grey except for copper and gold)
- **usually quite strong materials** (exceptions like the Alkali Metals which are untypically soft)
- **easily beaten into shape (malleable) or drawn into wire (ductile)**
- **solids sonorous**

Chemical Properties of metals

- **tend to form basic oxides that react with acids to form salts** (if the oxide is soluble in water it forms an alkali of pH > 7, universal indicator blue or violet)
- **most react with acids to form a salt and hydrogen**

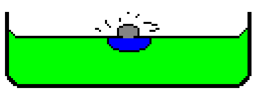
Typical Properties of Non-metallic Elements

Physical properties of non-metals

- **usually low melting points and boiling points and so can be gases, liquids or solids** (exceptions like silicon, and carbon as diamond or graphite)
- **poor conductors of heat and electricity** (exceptions like carbon in the form of graphite)
- **generally low density**
- **appearance - dull if solid**
- **usually weak materials eg soft or brittle solids** (exceptions like silicon, and carbon as diamond, which are very hard and strong)
- **if solid, not easily beaten into shape or drawn into wire, tend to be too brittle**
- **solids not usually sonorous**

Chemical properties of non-metals

- **form acidic oxides when burned in air or oxygen, these react with alkalis to form salts, if soluble in water they form acid solutions of pH <7, universal indicator yellow-orange-red**
- **they do not usually react with acids**



Typical Properties of Groups

- The very reactive **Group 1 The Alkali Metals** have low density (some float on water).

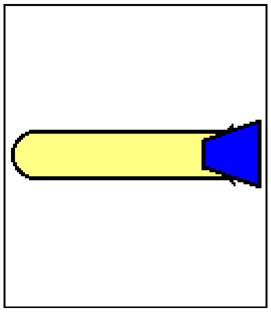
- They readily react with non-metals to form **ionic compounds** eg NaCl or $\mathbf{Na^+Cl^-}$, Li_2O or $\mathbf{(Li^+)_2O^{2-}}$.
- These are colourless crystals or white solids, soluble in water to give colourless solutions (usually pH 7 for their salts, pH 13-14 for the oxides because MOH alkali formed).
- The metals react rapidly, maybe violently, with water to form alkaline hydroxides and hydrogen gas.
- Alkali metal atoms have one outer electron, which is readily lost to form a stable single positive ion $\mathbf{M^+}$.
- Down the group, the metals get more reactive, and the melting points and boiling points decrease.

- **Group 2 are the 2nd group of metals** (sometimes called "Alkaline Earth Metals").
- They are not quite so reactive as the Alkali Metals for the same period.
- They have two outer electrons and readily lose them to form the $\mathbf{M^{2+}}$ ion.
- This ion occurs in the **ionic compounds** they readily form with non-metals like the Group 7 Halogens or oxygen and sulphur from Group 6 eg $\mathbf{MgCl_2}$ or **CaO**.

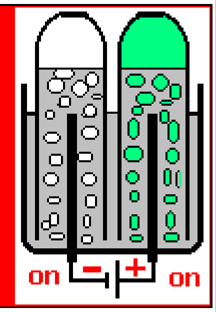
- **Group 3** contains the metal **Aluminium**.
- **Group 4** contains the non-metal **carbon** - which forms lots of compounds with hydrogen formed in oil
- **Group 5** contains the non-metal **nitrogen** - important element in natural and manmade artificial fertilisers. Nitrogen forms 79% ($^4/_5$th's) of air.

- **Group 6 are a Group of non-metallic elements**, the first 2 are **O oxygen** and **S sulphur**.
- They have 6 outer electrons and readily gain 2 electrons to form an $\mathbf{X^{2-}}$ ion in the **ionic compounds** they form with metallic elements eg in **MgO** and $\mathbf{Na_2S}$ or $\mathbf{Mg^{2+}O^{2-}}$ and $\mathbf{(Na^+)_2S^{2-}}$.
- They form **covalent small molecule compounds** with other non-metallic elements eg $\mathbf{H_2O}$ or $\mathbf{CS_2}$.
- The top element in the group is **oxygen, a most important element**.
  - Made by green plants in photosynthesis.
  - Consumed in the reverse process of respiration.

- Pure oxygen is obtained from the fractional distillation of liquified air,
- **Oxygen is used in**:
  - oxy-acetylene burners to produce a much hotter and intense flame for 'cutting' and welding metal,
  - oxygen 'tanks' in hospitals for respiratory problems,
  - oxidant gas for burning rocket fuel.



The Group 7 Halogens



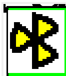
The Halogens are typical non-metals and form the **7th Group in the Periodic Table** 'Halogens' means **'salt formers'** and the most common compound is **sodium chloride** which is found from natural evaporation as huge deposits of 'rock salt' or the even more abundent 'sea salt' in the seas and oceans.

Physical features and important trends down the Group with increasing atomic number

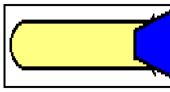
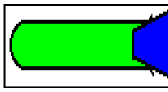
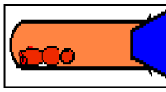
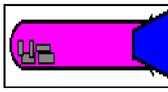
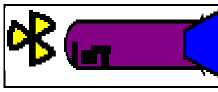
- **typical non-metals** with relatively **low melting points and boiling points**.
- the **melting points and boiling increase** steadily down the group (so the change in state at room temperature from gas => liquid => solid), this is because the intermolecular attractive forces increase with increasing size of atom or molecule.
- they are all **coloured non-metallic elements.**
- the **colour** of the halogen **gets darker** down the group.
- they are all **poor conductors of heat and electricity** - typical of non-metals.
- when **solid** they are **brittle and crumbly** eg **iodine**.
- the **size of the atom gets bigger** as more inner electron shells are filled going down from one period to another.

Chemical features, similarities and reactivity trend ...

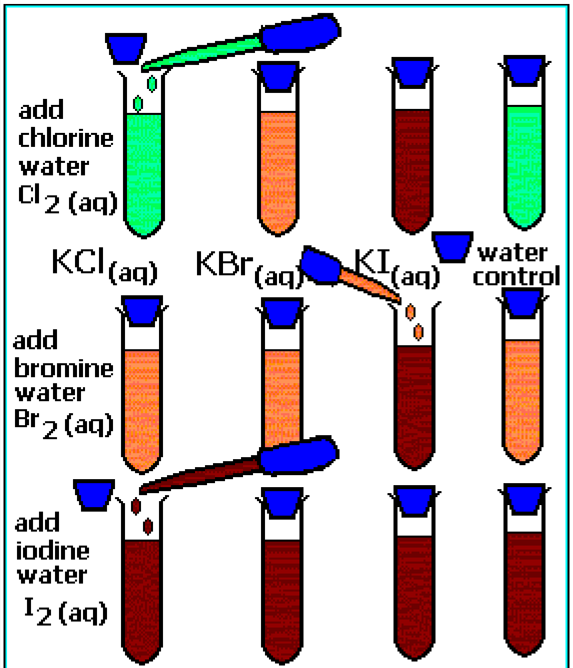
- The atoms all have **7 outer electrons,** this **outer electron similarity**, as with any Group in the Periodic Table, **makes them have very similar chemical properties** eg
  - they **form singly charged negative ions eg chloride Cl^-** because they are one electron short of a noble gas electron structure. They gain one negative electron (reduction) to be stable and this gives a surplus electric charge of -1. These ions are called the **halide ions**, two others you will encounter are called the **bromide Br^-** and **iodide I^-** ions.
  - they form **ionic compounds** with metals eg **sodium chloride Na^+Cl^-**
  - they **form covalent compounds with non-metals** and with themselves. The **bonding in the molecule** involves single covalent bonds eg **hydrogen chloride HCl** or **H-Cl**
- the elements all exist as **X_2 or X-X, diatomic molecules** where **X** represents the halogen atom.
- the **reactivity decreases down the group**
- they are all **TOXIC elements**.



- **Astatine is very radioactive**, so difficult to study BUT its properties can be predicted using the principles of the Periodic Table and the Halogen Group trends!

Selected Properties of the Group 7 Halogens (more AS-A2 data)						
Symbol and Name	**Atomic Number**	**Electron arrangement**	**State and colour at room temperature, colour of vapour when heated**	**Melting point**	**Boiling point**	**atom radius nm**
F Fluorine	9	2.7	 pale yellow gas	-220°C, 53K	-188°C 85K	0.072
Cl Chlorine	17	2.8.7	 green gas	-102°C, 173K	-34°C, 239K	0.099
Br Bromine	35	2.8.18.7	 dark red liquid, brown vapour	-7°C, 266K	59°C, 332K	0.114
I Iodine	53	2.8.18.18.7	 dark crumbly solid, purple vapour	114°C, 387K	184°C, 457K	0.133
At Astatine	85	2.8.18.32.18.7	 black solid, dark vapour	302°C 575K	380°C 653K	0.140

The Reactivity Order and Displacement Reactions



Chlorine water, bromine water and iodine water are added in turn to aqueous solutions of the salts potassium chloride (KCl), potassium bromide (KBr) and potassium iodide (KI). Three combinations produce a reaction (and 3 don't!). **You can get the observations from the diagrams! A darkening effect compared to a water blank confirms a displacement reaction has happened**. Chlorine **displaces** bromine from potassium bromide and iodine from potassium iodide. Bromine only displaces iodine from potassium iodide but iodine displaces non of the other two. On the basis that **the most reactive element displaces a least reactive element** the reactivity order must be **chlorine > bromine > iodine.**

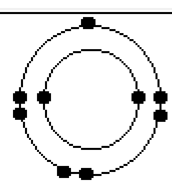
Equations: eg

- **chlorine + potassium bromide ==> potassium chloride + bromine**
- **$Cl_{2(aq)}$ + $2KBr_{(aq)}$ ==> $2KCl_{(aq)}$ + $Br_{2(aq)}$**
- **ionically the equations are written ...**

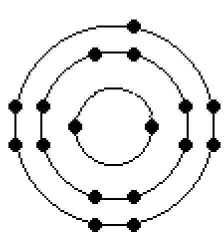
$Cl_{2(aq)} + 2Br^{-}_{(aq)} ==> 2Cl^{-}_{(aq)} + Br_{2(aq)}$

- the other 2 possible reaction equations are similar to the above example

- **REDOX: Oxidation-Reduction Theory**
- the **halogen molecule** is the electron acceptor (oxidising agent) and is **reduced by electron gain** to form a halide ion
- the **halide ion** is the electron donor (reducing agent) and is **oxidised by electron loss** to form a halogen molecule



F [2.7] + e^- ==> F^- [2.8]$^-$



Cl [2.8.7] + e^- ==> Cl^-

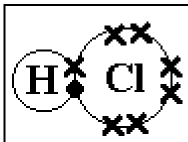
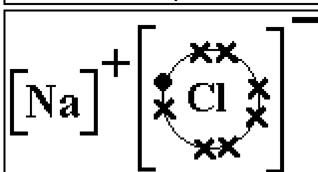
Explaining the Reactivity Trend of the Group 7 Halogens

- when a halogen atom reacts, it gains an electron to form a singly negative charged ion eg Cl + e^- ==> Cl^- which has a stable noble gas electron structure. (2.8.7 ==> 2.8.8)
- **as you go down the group** from one element down to the next .. F => Cl => Br => I
- the atomic radius gets bigger due to an extra filled electron shell
- the outer electrons are further and further from the nucleus and are also shielded by the extra full electron shell of negative electron charge
- therefore the outer electrons are less and less strongly attracted by the positive nucleus as would be any 'incoming' electrons to form a halide ion (or shared to

$[2.8.8]^-$ **Br $[2.8.18.7]$ + e^- => Br^- $[2.8.18.8]^-$ etc.**	form a covalent bond) • this combination of factors means to attract an 8th outer electron is more and more difficult, **so the element is less reactive as you go down the group**, ie less able to form the X^- ion with increase in atomic number

Other Reactions of the Halogens

note: fluorine forms fluorides, chlorine chlorides and iodine iodides

with hydrogen H_2	 Halogens readily combine with hydrogen to form the **hydrogen halides** which are colourless **gaseous covalent molecules.** eg hydrogen + chlorine ==> hydrogen chloride **$H_{2(g)}$ + $Cl_{2(g)}$ ==> 2HCl(g)** The **hydrogen halides dissolve in water** to form **very strong acids** with solutions of **pH1** eg hydrogen chloride forms **hydrochloric acid** in water **HCl(aq)** or **H^+Cl^- (aq)** because they are fully ionised in aqueous solution even though the original hydrogen halides were covalent! An acid is a substance that forms **H^+** ions in water. Bromine forms hydrogen bromide gas **$HBr_{(g)}$**, which dissolved in water forms hydrobromic acid $HBr_{(aq)}$. Iodine forms hydrogen iodide gas **$HI_{(g)}$**, which dissolved in water forms hydriodic acid $HI_{(aq)}$. Note the **group formula** pattern.
with Group 1 Alkali Metals Li Na K etc.	 Alkali metals burn very exothermically and vigorously when heated in chlorine to form **colourless crystalline ionic salts** eg **NaCl** or Na^+Cl^-. This is a very expensive way to make salt! Its much cheaper to produce it by evaporating sea water! eg **sodium + chlorine ==> sodium chloride** **$2Na_{(s)}$ + $Cl_{2(g)}$ ==> $2NaCl_{(s)}$** The sodium chloride is soluble in water to give a neutral solution pH 7, universal indicator is green. The salt is a **typical ionic compound** ie a brittle solid with a high melting point. Similarly potassium and bromine form potassium bromide **KBr**, or lithium and iodine form lithium iodide **LiI**. Again note the **group formula** pattern.
with other metals	If aluminium or iron is heated strongly in a stream of chlorine (or plunge the hot metal into a gas jar of chlorine carefully in a fume cupboard) the solid chloride is formed aluminium + chlorine ==> aluminium chloride$_{(white)}$: $2Al_{(s)}$ + $3Cl_{2(g)}$ ==> $2AlCl_{3(s)}$ iron + chlorine ==> iron(III) chloride$_{(brown)}$: $2Fe_{(s)}$ + $3Cl_{2(g)}$ ==> $2FeCl_{3(s)}$ If the iron is repeated with bromine the reaction is less vigorous, with iodine there is little reaction, these also illustrate the halogen reactivity series.

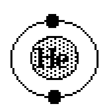
The Uses of Chlorine and other halogens and their compounds

 **CHLORINE**	All the Halogens are potentially harmful substances and **chlorine in particular is highly toxic**. It is highly dangerous to ingest halogens or breathe in any halogen gas or vapour. **Chlorine** is used to **kill bacteria** and so **sterilise water for domestic supply** or in **in swimming pools**. Organic phenolic chlorine compounds are used in **disinfectants like 'dettol' or 'TCP'** and organic chlorine compounds are used as **pesticides**. **Chlorine** is used in making **CFC refrigerant gases/liquids** but their production and use are being reduced. They break down in the upper atmosphere and the chlorine atoms catalyse the destruction of ozone O_3 which absorbs harmful uv radiation.
 VERY TOXIC TOO!	The sodium hydroxide and chlorine can be chemically combined to make the **bleach, sodium chlorate(I) NaClO**. This is used in some domestic cleaning agents, it chemically 'scours' and chemically 'kills' germs!
 	**Chlorine** (from electrolysis NaCl) **and ethene** (from cracking oil fraction) are used to make a chemical called **chloroethene** (which used to be called vinyl chloride). The chloroethene can be polymerised to form **poly(chloroethene)** which is very tough hard wearing useful plastic (old name **PVC**, polyvinyl chloride).
	**HCl$_{(aq)}$** As described above, some of the hydrogen and chlorine from the electrolysis of sodium chloride solution are combined to form **hydrogen chloride gas**. This gas is dissolved in water to **manufacture hydrochloric acid**. This is an important acid used in the chemical industry to make chloride salts.
 **silver salts** **Ag^+X^-**	**Silver chloride (AgCl), silver bromide (AgBr) and silver iodide (AgI) are all sensitive to light ('photosensitive'**), and all three are used in the production of various types of **photographic film** used to detect visible light and beta and gamma radiation from radioactive materials. Each silver halide salt has a different sensitivity to light. When radiation hits the film the silver ions in the salt are reduced by electron gain to silver (**Ag^+** + e^- ==> **Ag**, the halide ion is oxidised to the halogen molecule $2X^-$ ==> X_2 + $2e^-$). AgI is the most sensitive and used in X-ray radiography, AgCl is the most sensitive and used in 'fast' film for cameras.
FLUORINE F_2, BROMINE Br_2 and IODINE I_2	Fluorine is used as fluoride salts in toothpaste or added to domestic water supplies to strengthen teeth enamel helping to minimise tooth decay. (eg potassium fluoride). Apart from its silver salt use in photography, bromine s used to manufacture organic pesticides and fungicides because of their **poisonous nature** and **flame inhibitor chemicals** for plastic products to reduce their flammability. Also used, as well as **iodine**, in car headlamps. Iodine is used in hospitals in the mild **antiseptic** solution '**tincture of iodine**'.

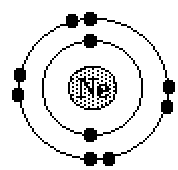
Introduction to the Group 0/8 Noble Gases

- **The "Noble Gases" are the last group in the Periodic Table ie they form the last elements at the end of a period.**
- **They are all non-metallic elements and all are colourless gases at room temperature and pressure with very low melting points and boiling points.**
- **They form 1% of air, and most of this is argon. All the noble gases, except radon, are separated by the fractional distillation of liquified air. Helium can also be obtained from natural gas wells where it has accumulated from radioactive decay (alpha particles become atoms of helium gas when they gain two electrons).**
- **They are very unreactive elements because the highest occupied electron level is complete, meaning they have a full shell of outer electrons! (see diagrams below). They have no 'wish' electronically to share electrons to form a covalent bond or to lose or gain electrons to form an ionic bond. In other words, they are electronically very stable.**
- **They exist as single atoms, that is they are monatomic He Ne Ar** etc. (NOT diatomic molecules as with many other gases - reasons given above).
- **Their very inertness is an important feature of their practical uses.**
- **Down the Group with increasing atomic number ...**
  - **the melting point and boiling point steadily increase**
  - **the density steadily increases**
  - **more likely to react and form a compound with very reactive elements like fluorine eg**
    - **Stable compounds of xenon are now known and synthesised BUT not before 1961!**

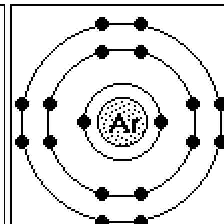
The first 3 Noble Gases, showing their electron arrangements with full very stable outer shells.



helium (2) 2



neon (10) 2.8

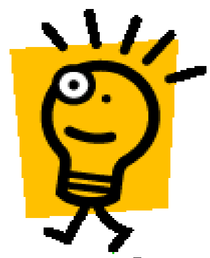
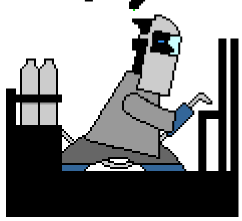
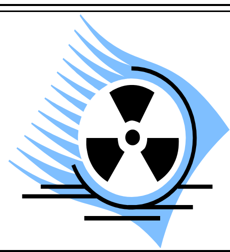


Ar (18) 2.8.8

Selected data on the Group 0/8 Noble Gases

Chemical symbol and name	Atomic number	Electron arrangement	Melting point	Boiling point	Atomic radius nm (10^{-9}m)
He helium	2	2	-270°C , 3K	-269°C , 4K	0.049
Ne neon	10	2.8	-249°C , 24K	-246°C , 27K	0.051
Ar argon	18	2.8.8	-189°C , 84K	-186°C , 87K	0.088
Kr krypton	36	2.8.18.8	-157°C , 116K	-152°C , 121K	0.103
Xe xenon	54	2.8.18.18.8	-112°C , 161K	-108°C , 165K	0.124
Rn radon	86	2.8.18.32.18.8	-71°C , 202K	-62°C , 211K	0.134

Uses of the the Group 0/8 Noble Gases

He helium		The gas is much less dense than air (lighter) and is used in balloons and 'airships'. Because of its inertness it doesn't burn in air UNLIKE hydrogen which used to be used in large balloons with 'flammable' consequences eg like the R101 airship disaster! Helium is also used in gas mixtures for deep-sea divers.
Ne neon		Neon gives out light when high voltage electricity is passed through it, so its used in glowing 'neon' advertising signs and fluorescent lights.
Ar argon	 	Argon, like all the Noble Gases is chemically inert. It used in filament bulbs because the metal filament will not burn in Argon and it reduces evaporation of the metal filament. It is also used to produce an inert atmosphere in high temperature metallurgical processes, eg in welding where it reduces brittle oxide formation reducing the weld quality. Its bubbles are used to stir mixtures in steel production. Argon is the cheapest to produce.
Kr krypton		Not used by superman! BUT is used in fluorescent bulbs, flash bulbs and laser beams.
Xe xenon		Good for winning scrabble games! AND also used in fluorescent bulbs, flash bulbs and lasers.
Rn radon		This has almost no uses, but does have dangers! Radio-isotopes of radon are produced by radioactive decay of heavy metals (eg uranium) in the ground. Can build up in cellars. Like all radio-isotopes it can cause cell damage (DNA) and ultimately cancer (see link below). However it is used in some forms of cancer treatment

Extra 'bits and bobs' on THE NOBLE GASES	
He helium, Ne neon, Ar argon, Kr krypton, Xe xenon, Rn radon	
% in Air by volume	**0.0005% He, 0.0018% Ne, 0.93% Ar, 0.0001% Kr, 0.00001% Xe, ?% Rn - impossible to be zero, but an extremely minute trace hopefully! (varies with local geology)**
Compounds of Noble Gases - yes they do exist!	**From the early 1960's compounds have been made, but only xenon compounds are stable and usually combined with oxygen and fluorine, which, not surprisingly, are the more reactive non-metals eg** **$Xe + 2F_2 => XeF_4$ (using Ni catalyst 60°C, easy if you know how!)**

Transition Metal Elements

- Cast iron is hard and used as man-hole covers. Steel is an alloy***** based on iron and used for car bodies. The ten horizontal elements **Sc to Zn** are called the 1st series of **Transition Metal Elements** eg iron and copper.
- These elements in the central blocks of the periodic table are typical metals - good conductors of heat and electricity and can be bent or hammered into shape (malleable) and they can be drawn into wire (ductile).
- However, compared to the group 1 Alkali Metals, they have higher melting points (except mercury - a liquid at room temperature); they are harder, tougher and stronger; they are much less reactive and so do not react (corrode) as quickly with oxygen or water.
- Most transition metals form coloured compounds (eg blue copper salt solutions) and are used in pottery glazes, stained glass and weathered copper roofs turn green!
- Many transition metals eg iron and platinum are used as catalysts. C
- *****alloy means a metal **mixed** with at least one other element.