Cameron Ramsay

Chemistry (Atoms, Bonds & Groups)

Atoms

- Properties of subatomic particles

Particle	Relative Mass	Relative Charge
Proton	1	+1
Neutron	1	0
Electron	1/2000	-1

- The atomic (proton) number is the number of protons in the nucleus
- The mass (nucleon) number is the number of protons and neutrons in the nucleus
- An **isotope** is an atom of a element with a different number of neutrons and therefore a different mass
- **Relative atomic mass (A_r)** is a weighted average of the **relative isotopic masses** of an element relative to the mass of Carbon-12
- **Relative molecular mass** is the sum of all the constituent atom's A_rs in a simple molecule eg. CO_2
- **Relative Formula mass** is the sum of all the constituent atom's A_rs in a giant structure eg. NaCl

Moles and Equations

- **Chemical Formulas**
    - Chemical formula tell you the number of each constituent atoms in a compound.
    - A number of standards are used when writing a chemical formula
        - Element's names are written using the element's symbol e.g. Oxygen is O, Magnesium is Mg
        - Subscript is used after an element's symbol to denote that there are multiple atoms of that element e.g. $CaCO_3$ - Calcium Carbonate, Cl_2 - Chlorine
- **Chemical Equations**
    - Chemical equations are used to show how an number of reactants, chemically react to become products.
    - The number of atoms of each element is always the same on both sides of the equation.
    - A coefficient is used before a chemical to denote that there are more than one of that molecule e.g. $2H_2O$ tells you that there are two molecules of water.
    - State symbols are added to show the state that the substances are, they are:

(s) Solid (l) Liquid (g) gaseous (aq) Aqueous

- An example of a chemical equation

$$Mg\ (s) + O_2\ (g) \rightarrow MgO\ (s)$$

This equation is not balanced, because there are two Oxygen atoms in the reactant side and only one in the products, to begin with make table of the number of each atom on each side.

$Mg\ (s) + O_2\ (g)$	$\rightarrow$	$MgO\ (s)$
1	Mg	1
2	O	1

To balance it we add coefficients to the Mg and the MgO (shown in red)

$2Mg\ (s) + O_2\ (g)$	$\rightarrow$	$2MgO\ (s)$
2 ~~1~~	Mg	~~1~~ 2
2	O	~~1~~ 2

The equation is now balanced

- Other balance equations
  - $2Al\ (s) + 3Cl_2\ (g) \rightarrow Al_2Cl_6\ (s)$
  - $4Na\ (s) + O_2\ (g) \rightarrow 2Na_2O\ (s)$

- **Number of moles**
  - The amount of a substance is the number of atoms is a sample, it is measured in moles
  - One mole of a substance contains 6.022×10^{23}, this is avogadro's constant.
  - 12 g of carbon-12 is one mole and therefore contains 6.022×10^{23} atoms
  - The three equations used to determine the number of moles of a substance are:

Form a Mass:	From a Gas's Volume	From a Solution
$n = \frac{Mass}{M}$	$n = \frac{Volume\ (in\ dm^3)}{24} = \frac{Volume\ (in\ cm^3)}{24000}$	$n = c \times V\ (in\ dm^3)$

- In the equation $2Al\ (s) + 3Cl_2\ (g) \rightarrow Al_2Cl_6\ (s)$, two atoms of Aluminium reactive with three Chlorine molecule, so 2 moles of Al reacts with 3 moles of Cl_2 or 1 mole of Al reacts with 1.5 moles of Aluminium. The coefficients show the ratio of reactants and products.
- Strategy:
  - Form a balanced symbol equation and the M_rs of the reactants and products
  - Work out number of moles of the given substance, divide this by the substance's coefficient so that you can (using the ratio of the coefficients) find the number of moles of the desired substance

        - Finally using one of the equations find the mass, volume or concentration of the substance using the equations
    - Example:

If 57.5g of sodium react with what mass of chlorine in the reaction:

$$2Na\ (s) + Cl_2\ (g) \rightarrow NaCl(s)$$

$$n = \frac{Mass}{M}$$

$$n = \frac{57.5}{23} = 2.5$$

∴ 2.5 moles of Na

$$2.5\ moles \div 2 = 1.25\ moles$$

$$1.25\ moles \times 1 = 1.25\ moles$$

∴ 1.25 moles of Cl_2

$$n = \frac{Mass}{M}$$

$$1.25 = \frac{Mass}{35.5 \times 2}$$

$$1.25 \times 71 = mass$$

Mass = 88.75 grams

- **Empirical Formula and Molar Mass**
    - Molar mass is the sum of all of the relative atomic masses in the substance (this can be an element, molecule or ion)
    - The amount of a substance n, mass m and molar mass M are all linked

$$n = \frac{m}{M}$$

    - The molecular formula is the exact number of each atom in a compound
    - The empirical formula is the simplest ratio of atoms in a compound
    - Empirical formula is used to describe giant structures e.g. NaCl, SiO_2
    - Calculating empirical formula
        - Make a column for each element, write the element's symbol and A_r at the top
        - Divide either the percentage or mass of the element in the compound by the A_r, this finds the molar ratio of the atoms
        - Divide these numbers by the smallest one, this will give the integer ratio of the atoms
        - Convert this ratio into a formula

- Example:

Find the empirical formula the compound consisting of nitrogen and oxygen containing 0.46g of nitrogen combined with 1.2g of oxygen.

N - 14.0

$$n = \frac{Mass}{M}$$

$$n = \frac{0.46g}{14}$$

n = 0.0329

$$\frac{0.0329}{0.0329}$$

= 1

O - 16.0

$$n = \frac{Mass}{M}$$

$$n = \frac{1.2g}{16}$$

n = 0.0750

$$\frac{0.0750}{0.0329}$$

= 2.28

≈ 2

∴ the empirical formula is NO_2

- From the empirical formula the molecular formula can be found
  - The M_r of the empirical formula must be found
  - The given M_r of the compound is divided by the empirical formula's M_r
  - The empirical formula is multiplied by the number obtained above.
- Example:

The M_r of the compound with empirical formula NO_2 is 92.0 find the molecular formula of the substance.

$$14 + (16 \times 2) = 46$$

$$\frac{92}{46} = 2$$

∴ The molecular formula is N_2O_4

- **Volume of a gas**
  - Avogadro's law states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules
  - At room temperature and pressure one mole of any gas occupies approximately 24.0 dm^3 (or 24000 cm^3)
  - The equation that links volume V and number of gas molecules n is:

$$n = \frac{V}{24dm^3}$$

  - Example:

What volume does 4 moles of chlorine gas occupy?

$$n = \frac{V}{24dm^3}$$

$$4 = \frac{V}{24dm^3}$$

$$V = 4 \times 24dm^3$$

$$V = 96dm^3$$

Acids and Bases

- An **Acid** is a proton donor, when dissolved in water an acid will release H^+ ions (protons)
- A **Base** is a proton acceptor, OH^- ions accept a proton to form water and NH_3 forms NH_4^+
- Common acids:
  - Hydrochloric Acid - HCl
  - Sulfuric Acid - H_2SO_4
  - Nitric Acid - HNO_3
- Common Bases are:
  - Metal Oxides
  - Metal Hydroxides
  - Ammonia
- Alkalis are bases that are soluble in water, they release OH^- ions when dissolved in water
- Common Alkalis:
  - Sodium Hydroxide - NaOH
  - Potassium Hydroxide - KOH
  - Aqueous Ammonia - NH_3
- When an H^+ ion reacts with a metal ion or NH_4^+ a salt forms.
- A salt is compound formed when an H^+ ion is replaced by a metal ion or Ammonium ion
- Acids reaction with:
  - Acid + Carbonate → Salt + Water + Carbon Dioxide
  - Acid + Base (metal oxide) → Salt + Water
  - Acid + Alkali (metal hydroxide) → Salt + Water
  - Acid + Ammonium → Salt (ammonium salt)
- A **hydrated** salt is a salt that contains water molecules
- An **anhydrous** salt is a salt that contains no water molecules
- **Water of crystallisation** is the water molecules that are part of the crystal structure of a compound

Redox

- Oxidation – A gain of oxygen OR a loss of electrons (OILRIG)
- Reduction – A loss of oxygen OR a gain of electrons (OILRIG)
- Redox reaction – a reaction in which oxidation and reduction occur
- Oxidising agent – A substance that oxidises another substance, and is itself reduced
- Reducing Agent – A substance that reduces another substance, and is itself reduced
- Oxidation State – tells you how many electrons each atom of an element has gained or lost forming a compound
- Oxidation States Rules
  - The oxidation number of an element is zero
  - The oxidation number generally equals the charge on the ions found in the periodic table (e.g. group 2 generally has +2)
  - The oxidation number of oxygen is -2
  - The oxidation number of hydrogen is +1 except in a metal hydride (hydrogen bonded with a metal) then it is -1

  - The oxidation number of the halogens is usually -1
  - The overall charge of a compound is equal to the sum of the oxidation numbers
- Roman numerals indicate the oxidation state of a substance
- Generally metals form positive ions, this is oxidation
- Generally non-metals form negative ions, this is reduction
- When a metal reacts with an acid the hydrogen is reduced and the metal is oxidised

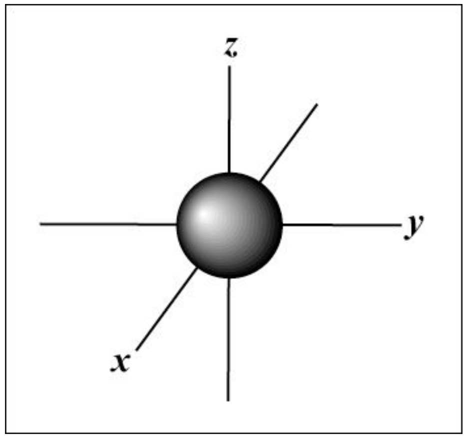
Electron Structure

- **First Ionisation Energy** is the energy required to remove one electron from each atom in one mole of a gas to form one mole of gaseous 1+ ions
- **Successive Ionisation Energies** are the energy required to remove one electron of each atom in one mole of a gas. eg. **second ionisation energy** is the energy required to remove one electron from each atom in one mole of 1+ gaseous ions, forming 2+ ions.
- **Change in Ionisation energies**
  - **Nuclear Charge** - The greater the nuclear charge, the greater the attraction between the nucleus and the outer electrons
  - **Atomic Radius** - The greater the atomic radius the smaller the nuclear attraction experienced by the outer electrons
  - **Electron Shielding** - Inner electrons repel the outer electrons, the more inner shells the more the electrons are repelled and therefore the lower the nuclear attraction between the outer electrons and the nucleus.
- Each time you remove an electron from a new shell the ionisation energy increases dramatically
- **The number of electrons in the first four shells**

n	Shell	Number of Electrons in Shell
1	1^{st} Shell	2
2	2^{nd} Shell	8
3	3^{rd} Shell	18
4	4^{th} Shell	32

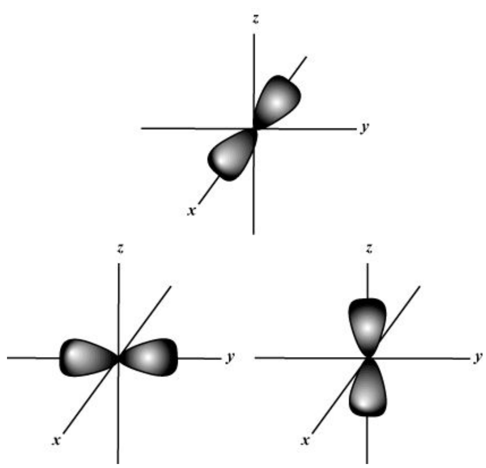
- An **Orbital** is a region that can hold up to two electrons with opposite spins
- The first three orbitals are **S, P** and **D**

- **S-orbitals**



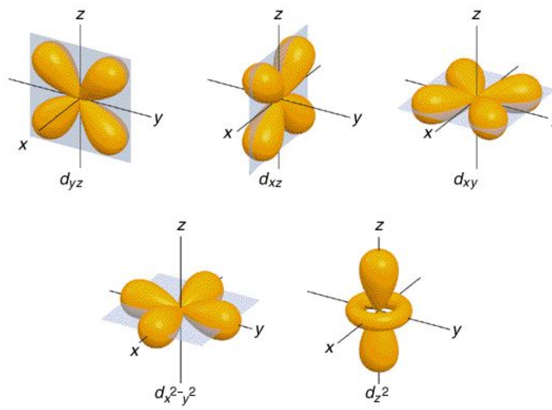
S Orbital shape

  - Each shell contains one S-orbital
  - Two electrons occupies the an S-orbital so each S subshell contains two electrons
- **P-orbitals**



P-orbital shape

  - Each shell from n = 2 upwards contains three P-orbitals
  - Each P-orbital contains two electrons so each P subshell contains 6 electrons
- **D-orbitals**



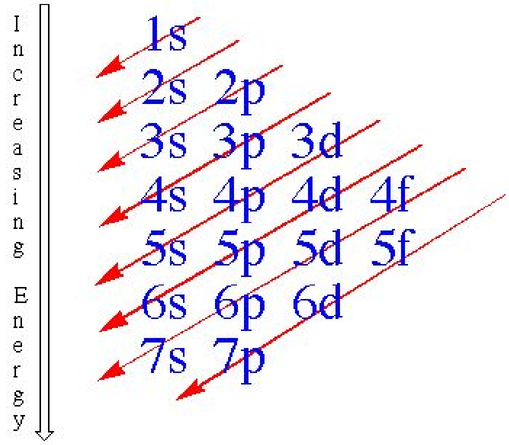
D-orbital shapes

- Each shell from n = 3 upwards contains five D-orbitals
- Each D-orbital contains two electrons so each D subshell contains 10 electrons

- **Energy levels**

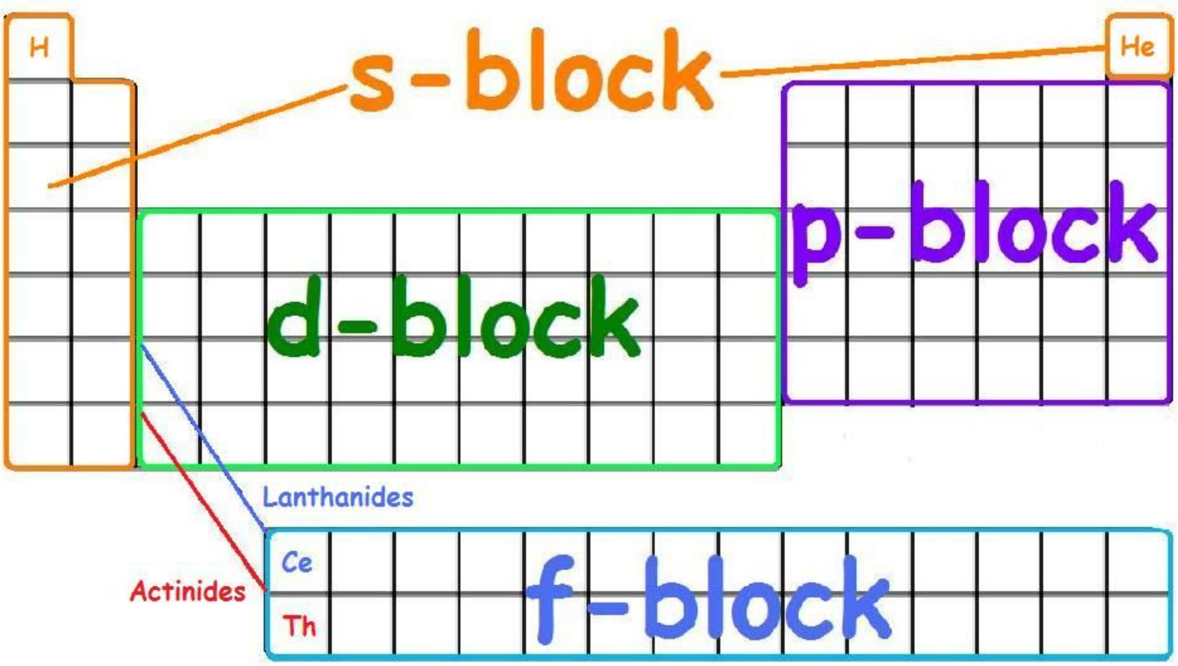
^					**4f**	________
|					**4d**	________
|					**4p**	________
|			**3d**	________		
Energy					**4s**	________
|			**3p**	________		
|			**3s**	________		
|		**2p** ________				
|		**2s** ________				
|	**1s** ________					

- **To help remember**



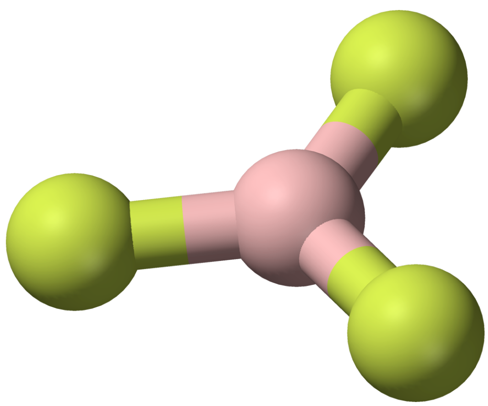
- An element's electron structure can be written in shorthand for example
  - O - $1s^2 2s^2 2p^4$
  - Br - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

- The outer orbital determines the block that the element is in.



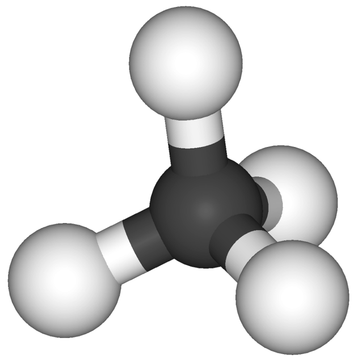
Bonding and Structure

- An **Ionic bond** is an electrostatic attraction between two oppositely charged ions, the metal atom give the nonmetal atom one or more electrons. Both atoms becoming charged
- An atom will gain or loose enough electrons to have a full outer shell.
- Common ions:
    - Nitrate Ions - NO_3^-
    - Carbonate Ions - CO_3^{2-}
    - Sulfate Ions - SO_4^{2-}
    - Ammonium Ions - NH_4^+
- A **Covalent bond** is a bond between two non-metal atoms, in which a pair of electrons is shared between the two atoms. The negative charge of the electrons pair attracts the positively charged nuclei.
- In a **Dative Bond** one atom supplies both of the electrons in the shared pair.
- The shape of a simple molecule is determined by the repulsion of electron pairs around the central atom.
- Lone electron pairs repel more than bonded pairs of electrons
- Molecule shapes
    - BF_3 - Trigonal Planar

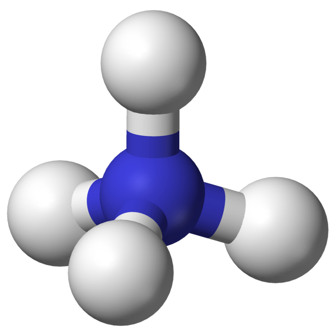


        - Bond angle: 120°
        - The electrons in the fluorine atoms repel each other so the bond angle is 120° as this is the angle that the bonds are the furthest distance away from each other.

- CH_4 - Tetrahedral

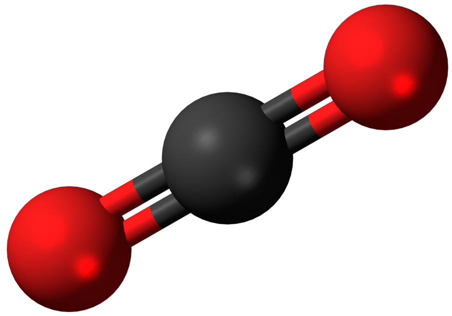


  - Bond angle of 109.5°
  - The electrons in the bond between the carbon and the hydrogen atoms repel each other so that the hydrogen atoms are as far from each other as possible.
- NH_4^+ - Tetrahedral



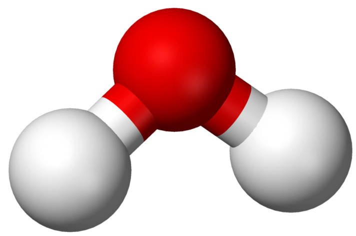
  - Bond angle of 109.5°
  - The electrons in the bond between the nitrogen and the hydrogen atoms repel each other so that the hydrogen atoms are as far from each other as possible.
- SF_6 - Octahedral
  - Bond angle of 90°
  - The electrons in the bond between the fluorine and the sulfur atoms repel each other so that the fluorine atoms are as far from each other as possible

- CO_2 - Linear



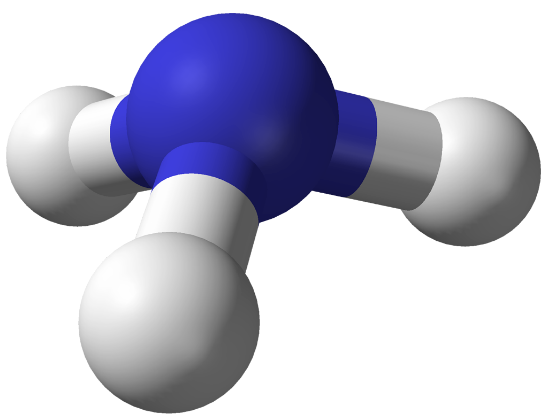
    - Bond angle of 180°
    - The bonded electrons in the carbon atom repel each other

- H_2O - Bent (non-linear)



    - Bond angle of 104.5°
    - The lone pairs of electrons in the oxygen atom repel the bonded electrons more than other bonded electrons, this reduces the angle of the bond.

- NH_3 - Pyramidal



    - Bond angle of 107°
    - The lone pair of electrons in nitrogen atoms repel the bonded electrons more than other bonded electrons do, this reduces the bond angle

- For each lone pair the bond angle is reduced by around 2.5°
- **Electronegativity** is the ability for an atoms to attract the bonding electrons in a covalent bond.

- The most electronegative element is fluorine and the least is Francium. Electronegativity increase as you move up and to the right in the periodic table.
- A permanent dipole is formed when the bonded atoms have different electronegativity, the pair of electrons will be closer to the more electronegative atom, so this side of the bond will be more negative and the other more positive.
- A polar molecule is one which has an overall dipole, when you take into account any dipole across the bonds.
- If a molecule has a permanent dipole it will form intermolecular bonds, holding the molecules together
- If there is no permanent dipole then the molecules will be held together by Van der Waals' forces. these are forces between temporary dipoles.
- Temporary dipoles form when there is a high concentration of electrons causes a charged area in the molecule
- O-H and N-H bonds have very dipoles, they form **hydrogen bonds** these are intermolecular bonds between the positive hydrogen and a lone pair of electrons on the oxygen or nitrogen
- Water has high melting and boiling points because the hydrogen bonds hold together the water molecules and so more energy is required to break them
- Ice is less dense than water because when solid the hydrogen bonds hold the water molecules apart in a lattice
- Metallic bonding is the electrostatic attraction of positive ions to delocalised electrons and occurs between metal atoms
- All ionic compounds exist as giant ionic lattices, in these each ion is surrounded by ions with an opposite charge. The opposite charges attract and the ions therefore are attracted to each other.
    - An example of this is NaCl
- A giant covalent lattice is a 3D structure in which all of the atoms are held together with strong covalent bonds
    - For example diamond or graphite
- A giant metallic lattice is a 3D structure of positive ions and delocalised electrons, bonded together with metallic bonds
- Simple molecular structures are 3D structures that contain covalently bonded molecules. The molecules are held together by weak intermolecular forces, while the constituent molecules contain atoms strongly bonded together
    - For example I_2 or ice.

<u>Periodicity</u>

- The periodic table is arranged
    - by increasing proton number
    - in periods showing repeating trends in physical and chemical properties
    - in groups having similar physical and chemical properties
- **Periodicity** is the repeating pattern of physical and chemical properties across the different periods

- Atoms in the same group have similar properties because they have similar outer electron structures
- There is a general increase in first ionisation energy across the period as the nuclear charge increases
- Theres is a decrease in first ionisation energy as you move down the group as the is an increase in atomic radius and electron shielding, outweighing the increasing nuclear charge

Group 2

- Redox reaction with oxygen

$$2X_{(s)} + O_{2(g)} \longrightarrow 2XO_{(s)}$$

- Redox reaction with water

$$X_{(s)} + 2H_2O_{(l)} \longrightarrow X(OH)_{2(aq\ or\ s)} + H_{2(g)}$$

- Reactivity increases down the group because
    - The first ionisation energy decreases because
        - Atomic size increases down the group
        - Electron Shielding increase down the group
        - Nuclear attraction increase however the other factors outweigh this increase
- Group 2 oxides react with water to form hydroxides, these solutions generally have a pH of 10-12

$$XO_{(s)} + H_2O_{(l)} \rightarrow X(OH)_{2\ (aq)}$$

- The solubility and alkalinity of group 2 hydroxides increase down the group
- Group 2 carbonates decompose with heat forming a metal oxide and CO_2, this is thermal decomposition
    - Group 2 carbonates become **harder** to decompose as you move **down** the group
- Group 2 elements are reducing agents, they react by losing electrons
- $Ca(OH)_2$ is used in agriculture to neutralise acidic soils
- $Mg(OH)_2$ is used in antacid tablets as it is alkali

Group 7

- Boiling points of group 7 elements increases as you move down the group
    - Because there are more electrons in the atom's shells, this means that there can be a large concentration of electrons so the van der Waal's forces are stronger
- Halogens exist as diatomic molecules
- Colours of halogen solutions in different solvents

Halogen	Water	Cyclohexane
Cl_2	Pale-green	pale-green
Br_2	Orange	Orange
I_2	Brown	Violet

- Chlorine is more reactive than Bromine so

$$Cl_{2\,(aq)} + 2Br^-_{(aq)} \rightarrow 2Cl^-_{(aq)} + Br_{2\,(aq)}$$

- Halogens become less able to form halide ions as you move down the group because
  - Atomic size increases so the electrons are attracted less to the nucleus
  - Electron shielding increases causing the electrons to be attracted less to the nucleus
  - Nuclear attraction increases but the other factors counter this increase
- **Disproportionation** is a reaction in which an element is both reduced and oxidised
- **Disproportionation of chlorine in water**
  - Chlorine is added to drinking water to kill bacteria, two acids are formed HCl and chloric(I) acid (HClO)
  - Chlorine is reduced and oxidised

$$Cl_{2\,(aq)} + H_2O_{(l)} \rightarrow HClO_{(aq)} + HCl_{(aq)}$$

- **Disproportionation of chlorine and aqueous sodium hydroxide**
  - This is the reaction used to form bleach
  - Occurs at room temperature
  - Chlorine is both oxidised and reduced

$$Cl_{2(aq)} + 2NaOH_{(aq)} \rightarrow NaCl_{(aq)} + NaClO_{(aq)} + H_2O_{(l)}$$

- Chlorine is water

Pros	**Cons**
Kills bacteria very effectively	Can react with hydrocarbons to form chlorinated hydrocarbons, suspect to cause cancer
	Chlorine gas is toxic

- Testing for halide ions
  - Halide substance dissolved in water, with silver nitrate ($AgNO_{3(aq)}$)
  - The silver ions react with the halide ions to form a silver halide precipitate
  - The colour of the silver halide precipitates tell us the halide present
  - **If unsure of colour**
  - Add aqueous ammonia NH_3
  - The different silver halides have different solubilities with aqueous ammonia
- The test are:

Chlorine	$Ag^+_{(aq)} + Cl^- \rightarrow AgCl_{(s)}$	White precipitate	Soluble in dilute NH_3
Bromine	$Ag^+_{(aq)} + Br^- \rightarrow AgBr_{(s)}$	Cream precipitate	Soluble in conc NH_3
Iodine	$Ag^+_{(aq)} + I^- \rightarrow AgI_{(s)}$	Yellow precipitate	Insoluble in conc NH_3