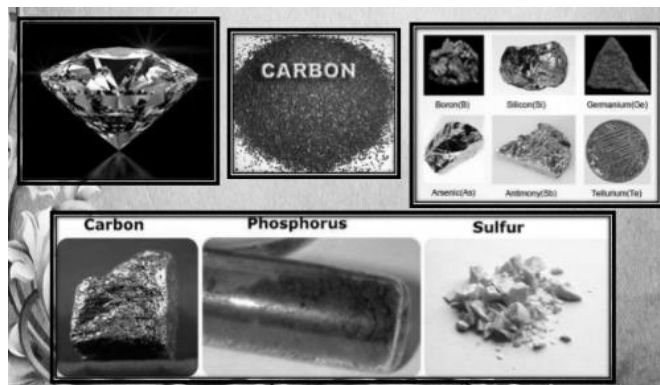


UNIT 12

NON-METALS AND THEIR COMPOUNDS

Student Version



Source: www.slideshare.net

Contents

- Introduction to Non-Metals
- The Noble Gases
- The Halogens and their Compounds
- Hydrogen, Oxygen, Oxides and Hydroxides
- Nitrogen and its Compounds
- Sulphur and its Compounds
- Carbon and its Compounds
- Rocks

Key words: noble gas, halogen, halide, halogen displacement, upward delivery, downward delivery, acidic oxide, basic oxide, amphoteric oxide, oxo-acid, Haber process, thermal decomposition, Contact process, peat, sub-bituminous coal, lignite, anthracite, coke, destructive distillation, coal tar, coal oil, coal gas, synthesis gas, mineral, rock, igneous, sedimentary, metamorphic, weathering

Units which must be completed before this unit can be attempted:

- Unit 1 – Atoms and the Periodic Table
- Unit 2 – Particles, Bonding and Structures
- Unit 3 – Amount of Substance and Measurement
- Unit 4 – Introduction to Physical Chemistry
- Unit 5 – Acids, Bases and Salts
- Unit 6 – Redox Reactions
- Unit 8 – Solubility and Precipitation Reactions
- Unit 9 – Metals and Their Compounds

Estimated Teaching Time: 15 hours

Lesson 1 – What are non-metals and what are some of their properties?

a) Introduction to non-metals



Summary Activity 1.1: What can you remember about Non-Metals?

- What is a non-metal?
- Give three examples of non-metals
- Describe the structure and bonding in diamond
- Describe the structure and bonding in chlorine

- **Non-metals** are elements which do not contain metallic bonds; the vast majority of non-metals contain covalent bonding; a small number (the noble gases) consist of free atoms; some non-metals have simple molecular structures (eg H_2 , O_2 , N_2), with the different molecules held together by intermolecular forces; some have giant covalent structures (eg Si, B, C) in which the covalent bonds extend throughout the lattice
- The physical properties of non-metals depend largely on whether they have giant covalent or simple molecular structures:
 - those with giant structures have high melting points, because the atoms are held together by strong covalent bonds; they are generally brittle, because the atoms cannot be moved around without breaking the covalent bonds; they are generally poor conductors of electricity as they have no free electrons or ions (although some, like graphite, do have free electrons and are exceptions); they can be soft or hard, depending on whether the structure is 2D layered (like graphite) or 3D and not layered (like diamond)
 - those with simple molecular and simple atomic structures have low melting points and are often gases at room temperature, as the Van der Waal's forces holding the molecules together are usually weak; they are all soft, weak and non-conductors of electricity

b) The Noble Gases

- The **noble gases (He, Ne, Ar, Kr and Xe)** are the elements which occupy Group 0 of the Periodic Table; atoms of these elements have full outer shells; these elements exist as free gaseous atoms and are extremely unreactive; argon is the most abundant of these elements and makes up around 1% of the atmosphere
- Because first ionisation energies increase across a Period from left to right, noble gases have the highest first ionisation energies; this means that they hold on to their electrons strongly and cannot form cations; because their outer shells are full, they can also neither accept nor share electrons, and so they remain as free atoms
- The noble gases all have very low boiling points due to the very weak Van der Waal's forces between the atoms; on descending the Group from He to Xe the number of electrons per atom increases and so the Van der Waal's forces become stronger and the boiling point increases; He boils at 4 K (-269 °C) and Xe boils at 166 K (-107 °C)
- The density of the noble gases increases down the group as the mass of the atoms increase; He and Ne are less dense than air; Ar, Kr and Xe are more dense than air

Unit 12 – Non-Metals and their Compounds

- The uses of the noble gases are related to their unreactivity and in some cases their density:
 - Helium is unreactive and less dense than air; it is used in party balloons and in hot air balloons; it is used in deep sea diving oxygen tanks instead of nitrogen (nitrogen dissolves in the blood at high pressures which can be very dangerous)
 - Neon is mainly used in lighting, as when heated it gives a red-orange colour without burning
 - Argon is used as an inert atmosphere when a reaction requires the absence of oxygen; argon will not take part in the reaction



Test your knowledge 1.2: Group 0 – the noble gases

- (a) Explain why the noble gases are so unreactive
- (b) State and explain the trend in boiling points of the noble gases
- (c) State and explain the trend in densities of the noble gases
- (d) State a use for each of helium, neon and argon

c) The Halogens and their Compounds

- The **halogens** are the collective name given to the elements in group 7 of the Periodic Table; the four important halogens are fluorine, chlorine, bromine and iodine

(i) Appearance and Physical Properties of the Halogens

- Since each atom in this group has seven valence electrons, they have one available space in their outer shell and so tend to form diatomic molecules, eg F_2 , Cl_2 , Br_2 and I_2 ; they thus have simple molecular structures, with weak intermolecular forces between the molecules; they show all the typical properties of simple molecular structures; they have low melting and boiling points, are poor conductors of electricity and tend to be soft and weak
- The colour and appearance of the halogens depends on whether they are in pure elemental form, dissolved in water or dissolved in a non-polar solvent:

Halogen	In pure form	In non-polar solvents	In water
Fluorine	Pale yellow gas	(Reacts with solvents)	(Reacts with water)
Chlorine	Pale green gas	Pale green solution	Pale green solution
Bromine	Dark red liquid	Orange solution	Orange solution
Iodine	Grey solid	Purple solution	(Insoluble) but forms a brown solution if excess KI is present

Unit 12 – Non-Metals and their Compounds

- The melting and boiling points of the halogens increase steadily down the group; this is due to the increase in strength of the Van Der Waal's forces between the molecules, which results from the increasing number of electrons in the molecule and the increasing surface area of the molecule:

Halogen	Melting point /°C	Boiling point /°C
Fluorine	-220	-188
Chlorine	-101	-35
Bromine	-7	59
Iodine	114	184



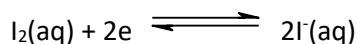
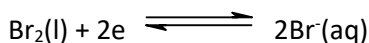
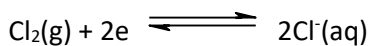
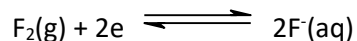
Test your knowledge 1.3: Group 7 – the halogens (physical properties)

- State the colour of pure iodine, iodine in cyclohexane and iodine in water
- State the colours of chlorine and bromine when dissolved in cyclohexane or water
- State and explain the trend in boiling points of the halogens
- The final halogen is astatine; suggest whether it will be a solid, liquid or gas at room temperature and suggest a possible melting point and boiling point

Lesson 2 – What are the main chemical properties of halogens and halides (part 1)?

(ii) Chemical Properties of Halogens and Halides

- All the halogens are oxidising agents, as they can accept electrons and get reduced (see Unit 6 – Redox Reactions):



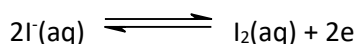
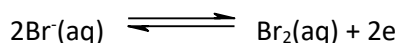
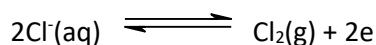
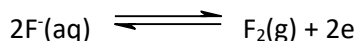
They are most commonly found in the -1 oxidation state, as X^- ions; these are known as **halide** ions

- Halide ions have no colour – they appear as colourless in aqueous solution and as white in the solid state
- Fluorine is the strongest oxidising agent, followed by chlorine; iodine is a mild oxidising agent; the fewer the number of shells in the atom, the closer the electrons can get to the nucleus and the less shielded the electrons are from the nucleus; the attraction of the electrons to the nucleus is thus stronger and the atom is more likely to accept electrons; thus the oxidising power of the halogens decreases down a group:

Fluorine	Chlorine	Bromine	Iodine
Very strong oxidising agent	Strong oxidising agent	Fairly strong oxidising agent	Mild oxidising agent

Unit 12 – Non-Metals and their Compounds

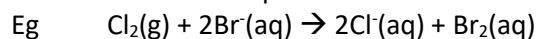
- The halide ions are reducing agents, as they can lose electrons and get oxidised (see Unit 6 – Redox Reactions):



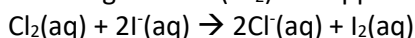
Iodide ions are the most reducing, followed by bromide ions; fluoride ions have no significant reducing properties; as the number of shells in the ion increases, there is more shielding of the nucleus and the outer electrons become less strongly held; these electrons are thus lost more easily and the halide ion is more readily oxidised; thus the reducing power of the halides increases down a group:

Fluoride	Chloride	Bromide	iodide
Very poor reducing agent	Poor reducing agent	Fairly poor reducing agent	Fairly good reducing agent

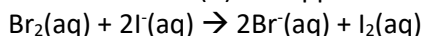
- The displacement reactions of halogens with halide ions provides a clear illustration of the trends in oxidizing properties of the halogens and the trends in reducing properties of the halide ions in aqueous solution; the more reactive halogens (ie the strongest oxidising agents) will displace the more reactive halides (ie the strongest reducing agents) from solutions of their ions:
 - chlorine will displace bromide and iodide ions from solution
 - bromine will displace iodide ions from solution, but not chloride ions
 - iodine cannot displace either bromide or chloride ions from solution



An orange colour (Br_2) will appear in solution on adding chlorine to bromide ions



A brown colour (I_2) will appear in solution on adding chlorine to iodide ions



A brown colour (I_2) will appear in solution after adding bromine to iodide ions



Test your knowledge 2.1: Group 7 – redox properties of halogens and halides

- State and explain the trend in oxidising power of the halogens
- State and explain the trend in reducing power of the halides
- State what you would observe when bromine is added to a solution of sodium iodide; write an equation for the reaction taking place and explain why this reaction takes place
- State what you would observe when bromine is added to a solution of sodium chloride; explain your observations

Unit 12 – Non-Metals and their Compounds



Practical 2.2: Halogen displacement reactions

You will have access to bottles containing chlorine water, bromine water and iodine water, bottles containing solutions of KCl, KBr and KI and a bottle of cyclohexane

- 1) Pour 2 cm³ of KCl(aq) into two separate test tubes
- 2) Add 2 cm³ of bromine water to one of the test tubes and 2 cm³ of iodine water to the other test tube and record your observations
- 3) Add 1 cm³ of cyclohexane to both test tubes and shake well; record your observations
- 4) Pour 2 cm³ of KBr(aq) into two separate test tubes
- 5) Add 2 cm³ of chlorine water to one of the test tubes and 2 cm³ of iodine water to the other test tube and record your observations
- 6) Add 1 cm³ of cyclohexane to both test tubes and shake well; record your observations
- 7) Pour 2 cm³ of KI(aq) into two separate test tubes
- 8) Add 2 cm³ of chlorine water to one of the test tubes and 2 cm³ of bromine water to the other test tube and record your observations
- 9) Add 1 cm³ of cyclohexane to both test tubes and shake well; record your observations
- 10) Summarise your observations in a table similar to the one below:

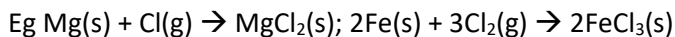
	Cl ₂		Br ₂		I ₂	
	water	cyclohexane	water	cyclohexane	water	cyclohexane
Cl ⁻						
Br ⁻						
I ⁻						

In which cases is the halogen present at the end of the reaction different from the halogen present at the start of the reaction? Write equations for the reactions which take place.

Why do some of the reactions not take place?

Lesson 3 – What are the main chemical properties of halogens and halides (part 2)?

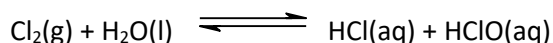
- Chlorine will combine directly with most metals to make metal chlorides; the chlorine is reduced (from 0 to -1) and the metal is oxidised:



Bromine and iodine can react with some of the more reactive metals but the reactions tend to be slower and less vigorous

Unit 12 – Non-Metals and their Compounds

- Chlorine is slightly soluble in water and also reacts slightly with water:



- The HCl produced makes the solution acidic, so chlorine is regarded as an acidic gas
 - note that the chlorine is simultaneously oxidised and reduced in this reaction - from 0 to -1 (in HCl) and +1 (in HClO)
 - HClO, or chloric (I) acid, is a mild oxidising agent and effective at killing bacteria without being harmful to humans; for this reason a small amount of chlorine dissolved in water will sterilise the water and chlorine is widely used in swimming pools and in water treatment
 - chlorine is toxic to humans in anything other than very small doses, so care must be taken not to over-chlorinate the water supply
 - This reaction can be used to test for chlorine; add damp blue litmus paper to the gas; the litmus paper will first turn red and then white (the chlorine reacts with water to produce HCl, which turns blue litmus red, and HClO, which is a bleaching agent and turns the litmus paper white)
- Chlorine reacts readily with dilute alkali:
$$\text{Cl}_2(\text{g}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{NaClO}(\text{aq}) + \text{H}_2\text{O}(\text{l})$$
 - Again; the chlorine is simultaneously oxidised and reduced in this reaction - from 0 to -1 (in HCl) and +1 (in NaClO)
 - the ion ClO^- is known as the chlorate (I) ion; it is an important oxidising agent and is the active ingredient in domestic bleach; it is most commonly found as **NaClO (sodium chlorate (I))**
$$\text{ClO}^- + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{Cl}^- + \text{H}_2\text{O}$$

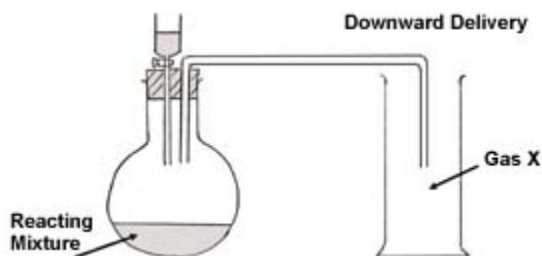


Test your knowledge 3.1: Group 7 – reactions of halogens with metals, water and alkalis

- Magnesium reacts with both chlorine and bromine; write equations for the reactions taking place
- Explain why magnesium reacts faster with chlorine than with bromine
- Write an equation for the reaction of chlorine with water and hence explain why chlorine is added to swimming pools and drinking water
- Write an equation for the reaction of chlorine with dilute sodium hydroxide and explain why the product is useful
- Describe how the oxidation number of Cl changes during the reactions in (c) and (d)
- Describe a simple laboratory test for chlorine gas

Unit 12 – Non-Metals and their Compounds

- Solid samples of the halide ions react with concentrated sulphuric acid
 - concentrated sulphuric acid is both an acid and an oxidising agent
 - Br^- and I^- ions, which are stronger reducing agents, are oxidised by concentrated sulphuric acid and so a redox reaction takes place
$$2\text{H}_2\text{SO}_4(\text{l}) + 2\text{KBr}(\text{s}) \rightarrow \text{SO}_2(\text{g}) + \text{Br}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + \text{K}_2\text{SO}_4(\text{s})$$
this is a redox reaction; Br^- is oxidised to Br_2 ; Br_2 can be seen as an orange vapour; SO_2 can be tested for using dichromate paper (see Unit 6 – Redox Reactions)
$$5\text{H}_2\text{SO}_4(\text{l}) + 8\text{KI}(\text{s}) \rightarrow \text{H}_2\text{S}(\text{g}) + 4\text{I}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l}) + 4\text{K}_2\text{SO}_4(\text{s})$$
this is a redox reaction; I^- is oxidised to I_2 ; I_2 can be seen as a purple vapour
 - Cl^- is a much weaker reducing agent so is not oxidised by concentrated sulphuric acid; an acid-base reaction takes place and HCl is produced:
$$\text{H}_2\text{SO}_4(\text{l}) + 2\text{KCl}(\text{s}) \rightarrow \text{K}_2\text{SO}_4(\text{s}) + 2\text{HCl}(\text{g})$$
this is not a redox reaction; white fumes of HCl are seen
 - This is the standard laboratory method for the preparation of HCl gas; note that because HCl is more dense than air and highly soluble in water, it is collected by downward delivery:



Reacting Mixture: concentrated H_2SO_4 added to a solid sample of KCl
Gas X = HCl

www.saburchill.com

- The identity of the gas can be confirmed by its chemical test: the gas will turn blue litmus red and will give a white smoke in the presence of filter paper soaked in ammonia (see Unit 5 – Acids, Bases and Salts)



Test your knowledge 3.2: Group 7 – reactions of halides with concentrated sulphuric acid

- Write an equation for the reaction of potassium chloride with concentrated sulphuric acid
- Write an equation for the reaction of potassium bromide with concentrated sulphuric acid
- Write an equation for the reaction of potassium iodide with concentrated sulphuric acid
- Explain how the equations in (g), (h) and (i) illustrate the trend in reducing properties of the halides
- Explain how HCl is prepared in the laboratory and explain why it should be collected by downward delivery



Demonstration 3.3: Prepare HCl gas in the laboratory

(CAUTION – HCl gas is highly toxic – use a fume cupboard; concentrated H₂SO₄ is highly corrosive; wear gloves)

- 1) Set up the apparatus as shown in the diagram above (a round-bottomed flask, a dropping funnel, a delivery tube, a bung to fit the flask with holes both for a delivery tube and a dropping funnel, a gas jar)
- 2) Add around 5 cm³ of concentrated H₂SO₄ into the dropping funnel; add around 2 g of KCl into the round-bottomed flask
- 3) Gradually add the concentrated sulphuric acid to the potassium chloride; you will observe a gas being given off which will collect in the gas jar
- 4) After around 30 seconds; close the tap of the dropping funnel, remove the delivery tube from the gas jar and place a lid on the gas jar
- 5) Confirm that the gas is HCl by dipping briefly some damp blue litmus paper into the gas jar, and then a piece of filter paper soaked in NH₃)

Why is the HCl collected by downward delivery?

Why is the HCl not collected over water?

Write an equation for the reaction taking place which produces HCl

Write an equation for the reaction taking place on the damp litmus paper

Write an equation for the reaction taking place with NH₃



Demonstration 3.4: Compare the reactions of concentrated sulphuric acid with the different halide ions

(CAUTION – the reaction releases toxic gases – use a fume cupboard; concentrated H₂SO₄ is highly corrosive; wear gloves)

- 1) Take three test tubes; to the first, add 2 g of solid KCl; to the second, add 2 g of solid KBr; to the third, add 2 g of solid KI; place all three test tubes in a test tube rack in a fume cupboard
- 2) Using a dropping pipette, gradually add 2 cm³ of concentrated H₂SO₄; note the colour of any gases produced
- 3) Test the vapours produced with ammonia and acidified potassium dichromate paper

What is the colour of the vapour produced with KCl and what is this vapour?

What is the colour of the vapour produced with KBr and what is causing this colour?

What is the colour of the vapour produced with KI and what is causing this colour?

In which test tube is a redox reaction not taking place?

In the other test tubes, what is being oxidised and what is being reduced?

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=I5O5dYEdO4

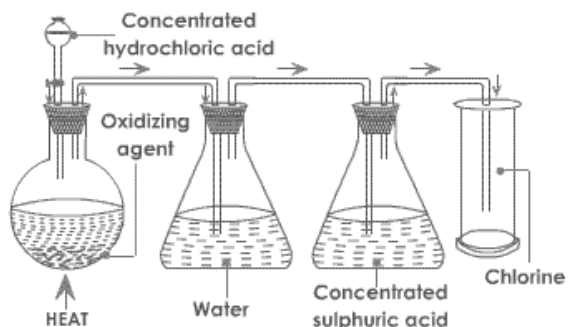
Lesson 4 – What are the main chemical properties of halogens and halides (part 3)?



Summary Activity 4.1: What can you remember about tests for halide ions and HCl gas?

- Describe a simple test for HCl
- Describe a simple test to identify chloride, bromide and iodide ions in solution
- Describe a simple test to identify chloride, bromide and iodide ions from a solid sample

- It is possible to oxidise chloride ions to chlorine in the presence of a very strong oxidising agent, such as MnO_2 ; the chlorine should be bubbled through water first (to remove HCl) and then through concentrated sulphuric acid (to remove water); the chlorine can then be collected by downward delivery (because it is more dense than air); $\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + 2\text{H}_2\text{O}$
The Mn is reduced from +4 to +2; the Cl^- is oxidised to Cl_2 (-1 to 0)



www.chem-guide.blogspot.com



Demonstration 4.2: Prepare Cl_2 gas in the laboratory

(caution – concentrated HCl and concentrated H_2SO_4 are very corrosive; wear gloves)

- 1) Set up the apparatus as shown in the diagram above (a round-bottomed flask, two conical flasks, a dropping funnel, three delivery tubes, three bungs to fit the flasks, each with two holes, a gas jar)
- 2) Add around 5 cm^3 of concentrated HCl into the dropping funnel; add around 5 g of MnO_2 into the round-bottomed flask; add around 50 cm^3 of water to the first conical flask and 50 cm^3 of concentrated sulphuric acid to the second conical flask
- 3) Gradually add the concentrated HCl to the MnO_2 ; you will observe a gas being given off which will eventually collect in the gas jar
- 4) When a reasonable quantity of gas has been collected in the gas jar, close the tap of the dropping funnel, remove the delivery tube from the gas jar and place a lid on the gas jar
- 5) Confirm that the gas is Cl_2 by dipping briefly some damp blue litmus paper into the gas jar

Write an equation for the reaction taking place

What is oxidised and what is reduced?

Why is the Cl_2 bubbled through first water and then concentrated H_2SO_4 ?

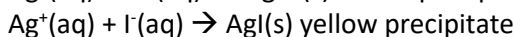
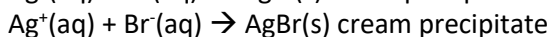
Why is the gas then collected using downward delivery?

Why does Cl_2 turn damp litmus paper red and then bleach it?

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=C2kf7OBbFsM

Unit 12 – Non-Metals and their Compounds

- Chloride, bromide and iodide ions all form precipitates with the silver (I) ion, Ag^+ ; each has its characteristic colour and these precipitation reactions can thus be used as chemical tests for these ions (see Unit 8 – Solubility and Precipitation Reactions; nitric acid is added before the silver ions to remove any hydroxide or carbonate ions which will also form a precipitate



(iii) Uses of Halogens and Halides

- Chlorine** is added to water as a sterilising agent because of its oxidising properties; **iodine**, a milder oxidising agent, is also a sterilising agent and is applied to wounds to prevent infection
- Sodium chlorate (I) (NaClO)** is the main ingredient in bleach; the ClO^- ion is a strong oxidising agent and kills bacteria: $\text{ClO}^- + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{Cl}^- + \text{H}_2\text{O}$
- Silver bromide (AgBr)** is used in conventional photography; it is applied to photographic film and decomposes on exposure to light, leaving silver: $2\text{AgBr}(\text{s}) \rightarrow 2\text{Ag}(\text{s}) + \text{Br}_2(\text{l})$; when the silver bromide is washed off, the silver remains, leaving a dark area on the parts of the film which were exposed to light; this is the “negative” of the photograph



Test your knowledge 4.3: Group 7 – other reactions of halogens and halides

- Describe the laboratory preparation of chlorine; state what type of reaction is taking place and explain how the chlorine is collected
- Describe a simple test for chloride, bromide and iodide ions in aqueous solution
- Give a use for chlorine, iodine, sodium chlorate (I) and silver bromide

Lesson 5 – What do I need to know about hydrogen and hydrogen ions?

d) Hydrogen, Oxygen, Oxides and Hydroxides



Summary Activity 5.1: What do you know about hydrogen?

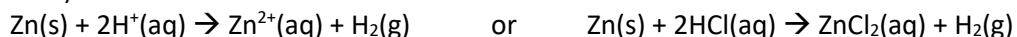
- Describe the structure and bonding in hydrogen gas
- Write an equation for a reaction between any metal and any acid. Which metals react with acids? What gas is produced in this reaction?
- Describe a simple test for H^+ ions in aqueous solution

(i) Hydrogen

- Hydrogen has a simple molecular structure; it exists as a diatomic gas (H_2); it is the most abundant element in the universe but it is not present in the atmosphere as it has a very low density and escapes quickly

Unit 12 – Non-Metals and their Compounds

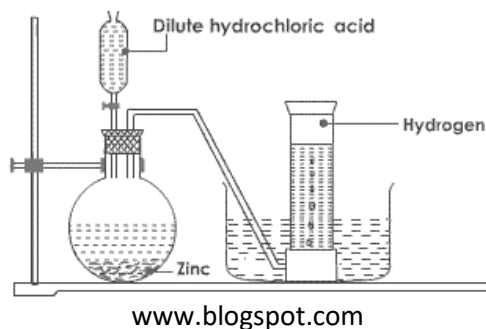
- Hydrogen is most conveniently prepared in the laboratory by reacting a fairly reactive metal (such as zinc) with acids to produce a salt and water (this reaction is also studied in Unit 6 – Redox Reactions as a way of preparing salts):



The metal is oxidised (from 0 to +2) and the H^+ in the acid is reduced from +1 to 0

This reaction works with any metal more reactive than hydrogen, but it is dangerous with very reactive metals

If the reaction is being used to prepare hydrogen, the following apparatus should be used:



Hydrogen is insoluble in water as it is non-polar; it can therefore be collected over water

- Hydrogen reacts rapidly with oxygen to form water; if a sample of hydrogen gas is exposed to a burning splint, it will give a “squeaky pop” sound when it burns; **this is the standard test for hydrogen gas**
- The flammability of hydrogen is the reason it is no longer used in hot air balloons



Demonstration 5.2: Prepare H_2 gas in the laboratory

Set up the apparatus as shown in the diagram above (a flask, a dropping funnel, a delivery tube, a bung to fit the flask with holes both for a delivery tube and a dropping funnel, a trough of water and a gas jar)

- Add around 10 cm^3 of dilute HCl into the dropping funnel; add around 20 zinc granules into the flask
- Gradually add the HCl to the zinc; you will observe a gas being given off which will collect in the gas jar
- Once the gas jar is almost full, slide a lid under the gas jar, remove it from the water; keep the lid of the gas jar under the jar
- Confirm that the gas is H_2 by lighting a splint, lifting the gas jar and placing the splint below the gas

Write an equation for the reaction taking place

What is oxidised and what is reduced?

Why can H_2 be collected over water?

Why should the lid of the gas jar always be placed at the bottom of the jar?

What is the test for hydrogen gas?

What reaction is taking place during this test?

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=K45ivcudgqY

Unit 12 – Non-Metals and their Compounds

(ii) Hydrogen Ions (H^+)

- Hydrogen can be oxidised and lose its (only) proton in aqueous solution to form a H^+ ion; the H^+ ion cannot exist in the solid state – it only exists in aqueous solution; its presence in solution makes solutions acidic (see Unit 5 – Acids, Bases and Salts)
- H^+ is usually formed when other acids dissolve in water: eg $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$
- Hydrogen ions turn blue litmus paper red; they will also give bubbles/fizzing if added to a solid or solution containing CO_3^{2-} or SO_3^{2-} ions (see Unit 5 – Acids, Bases and Salts)



Test your knowledge 5.3: Preparation and reactions of hydrogen

- (a) Describe the laboratory preparation of hydrogen; state what type of reaction is taking place and explain how the hydrogen is collected
- (b) Describe a simple test for hydrogen gas
- (c) Describe a simple test for hydrogen ions in aqueous solution

Lesson 6 – Why is water special?



Summary Activity 6.1: What can you remember about water?

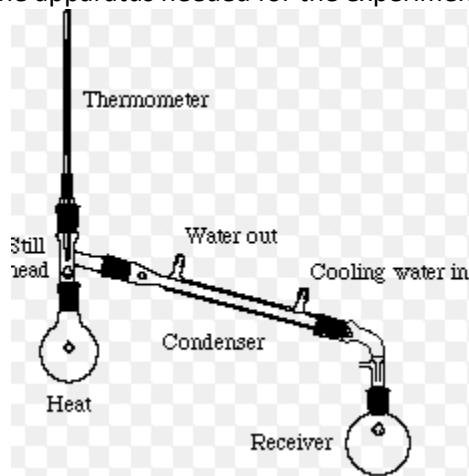
- Describe the structure and bonding in water
- Describe how you would measure the boiling point of water
- Explain why the boiling point of water is usually high
- Describe a simple test for the presence of hydroxide ions in solution

(iii) Water

- Water (H_2O) is the most common oxide of hydrogen; the other oxide of hydrogen is hydrogen peroxide (H_2O_2)

Unit 12 – Non-Metals and their Compounds

- Water has a simple molecular structure (see Unit 2 – Particles, Bonding and Structure), but despite being a small molecule with weak Van der Waal's forces, it has a relatively high melting point (273 K or 0 °C) and boiling point (373 K or 100 °C) meaning that it is a liquid at room temperature; its high melting point and boiling point is due to the presence of hydrogen bonding; the Celsius scale was created by setting the melting and boiling points of water at 0 °C and 100 °C respectively; an experiment to measure the boiling point of water was studied in Unit 2 – Particles, Bonding and Structure; the apparatus needed for the experiment is shown below:



www.chemconnections.org

- The hydrogen bonding in water is responsible for other unusual properties of water; the low density of ice is because the water molecules adopt a structure which maximises hydrogen bonding; the resulting structure is open and has a low density; water also has a very high surface tension, which is responsible for the noticeable meniscus in thin tubes and is the reason some insects can walk on water
- Hydrogen bonding is only possible in polar molecules; a simple electrostatic experiment can be used to demonstrate the polarity of water:



Demonstration 6.2: Demonstrate the polarity of water

- Inflate a balloon and rub it against an item of clothing (this makes it electrostatically charged)
- Turn a tap on and allow a steady stream of water to run
- Hold the balloon close to the stream of water

What happens?

Why does this happen?

Can you think of any other liquids which would give the same effect?

Can you think of any liquids which would not give this effect?

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=8RSDK17InUc

- The polar nature of water makes it a very good solvent; most ionic compounds are soluble in water because the polar water molecules are attracted to the ions; simple molecular compounds are also soluble in water if they are polar; see Unit 2 – Particles, Bonding and Structure and also Unit 8 – Solubility and Precipitation Reactions)

Unit 12 – Non-Metals and their Compounds

(iv) Hydroxide ions

- Water reacts with bases to produce solutions containing hydroxide ions: eg $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$ (see Unit 5 – Acids, Bases and Salts)
- Hydroxide ions turn red litmus blue; if warmed with ammonium chloride, hydroxide ions will give off ammonia gas which has a characteristic pungent smell (see Unit 5 – Acids, Bases and Salts)



Practical 6.3: Investigating the solubility of different substances in water

- 1) Place seven test tubes into a test tube rack and place 5 cm³ of water into each test tube
- 2) Into the first test tube, place a spatula of salt; into the second, place a spatula of sugar; into the third, a piece of chalk; into the fourth, a spatula of sand; into the fifth, a piece of wax; into the sixth, 2 cm³ of ethanol and into the seventh, 2 cm³ of paraffin
- 3) Shake each test tube well for a few minutes and then allow them to stand for a few minutes; note your observations

Which substances are soluble in water?

Can you explain the solubility of each substance in water in terms of its structure?

Lesson 7 – What else do I need to know about water, and what is hydrogen peroxide?

- The presence of dissolved substances in water can be tested in a number of ways:
 - in various precipitation reactions (see Unit 8 – Solubility and Precipitation Reactions); these reactions can confirm the presence of a number of ions
 - from indicators or pH meters (see Unit 5 – Acids, Bases and Salts); pure water has a pH of 7; if water contains dissolved acids or alkalis its pH will be lower or greater than 7
 - by testing whether it conducts electricity (see Unit 6 – Redox Reactions); pure water does not conduct electricity but water containing dissolved ions (electrolytes) does conduct electricity
 - by heating the water until most of it evaporates (see Unit 8 – Solubility and Precipitation Reactions); dissolved substances in water will often crystallise out as the water evaporates
- Water can form complex ions with d-block metal ions such as Cu^{2+} , which forms a blue complex ion with water; anhydrous copper sulphate turns blue when water is added and this is a test for water (see Unit 9 – Metals and their Compounds)
- Water is the most useful substance on earth; it is the basis for human life as the human body relies on water to transport minerals and nutrients around the body and as a solvent for chemical reactions in cells; water is also used for washing; water is used as an industrial coolant, especially in power stations; water is also used as a raw material for making other useful substances, such as ethanol

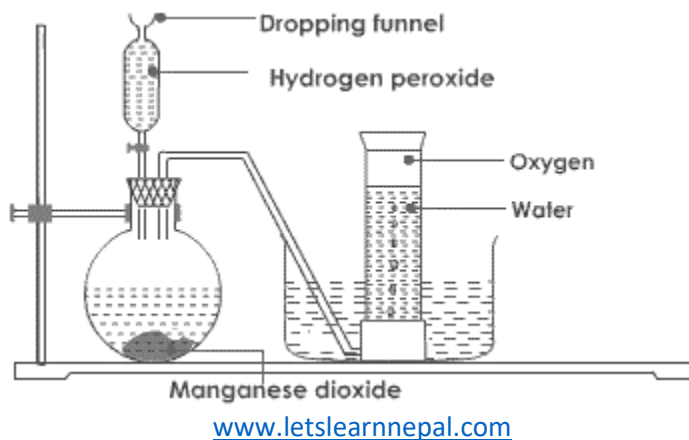
Unit 12 – Non-Metals and their Compounds

- Many places on earth provide a continuous supply of water which is pure and clean enough to drink; most water sources need to be treated before the water is treatable; the main steps in the treatment of water are as follows:
 - Coagulation: positively charged ions are added to the water; this results in many dirt particles and many other dissolved substances to coagulate and form larger particles
 - The larger particles settle to the bottom of the water supply and can be separated
 - The water remaining is filtered to remove the smaller particles
 - A disinfectant is then added to kill any dissolved bacteria which may be harmful

(v) Hydrogen Peroxide

- **Hydrogen peroxide (H₂O₂)** is a less common oxide of hydrogen; it contains O in an oxidation state of -1; all compounds containing O in an oxidation state of -1 are called peroxides
- Hydrogen peroxide is a useful oxidising agent known for its bleaching properties; it is especially used in the bleaching of hair
- Hydrogen peroxide decomposes in the presence of a catalyst (MnO₂) to give water and oxygen:
$$2\text{H}_2\text{O}_2(\text{l}) \rightarrow 2\text{H}_2\text{O} + \text{O}_2(\text{g})$$

In this reaction the O is both oxidised (from -1 to 0, in O₂) and reduced (from -1 to -2, in H₂O)
This reaction is a useful way of making oxygen in the laboratory; oxygen has limited solubility in water so can be collected over water



- Oxygen is a necessary component of combustion, and so will re-light a glowing splint; this is the standard test for oxygen gas



Demonstration 7.1: Prepare O₂ gas in the laboratory

Set up the apparatus as shown in the diagram above (a flask, a dropping funnel, a delivery tube, a bung to fit the flask with holes both for a delivery tube and a dropping funnel, a trough of water and a gas jar)

- 1) Add around 100 cm³ of water and one spatula of MnO₂ into the flask; add around 25 cm³ of concentrated H₂O₂ (2 moldm⁻³) to the dropping funnel
- 2) Gradually add the H₂O₂ to the flask; you will observe a gas being given off which will collect in the gas jar
- 3) Once the gas jar is almost full, slide a lid under the gas jar and remove it from the water
- 4) Confirm that the gas is O₂ by lowering a glowing splint into the jar

Write an equation for the reaction taking place

What is oxidised and what is reduced?

Why can O₂ be collected over water?

What is the test for oxygen gas?

What reaction is taking place during this test?

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=nkeniDKGs6Q



Test your knowledge 7.2: Water and hydrogen peroxide

- (a) Give three unusual properties of water and explain why water has these properties
- (b) Name two substances which dissolve in water and two which do not
- (c) State four ways in which you could detect whether there are any dissolved substances in water
- (d) Describe a simple chemical test for water
- (e) Describe how water is treated so that it is safe to drink
- (f) Draw the structure of hydrogen peroxide
- (g) Write an equation for the decomposition of hydrogen peroxide and explain why this reaction is useful

Lesson 8 – What do I need to know about oxygen and oxides?



Summary Activity 8.1: What can you remember about oxygen and oxides?

- Describe the structure and bonding in oxygen
- Describe the laboratory preparation of oxygen and explain why the gas can be collected in this way
- Describe a simple chemical test for oxygen gas
- Identify two basic oxides and write equations for their reactions with HCl
- Identify two acidic oxides and write equations for their reactions with NaOH
- Identify one amphoteric oxide and write equations for its reactions with HCl and NaOH

Unit 12 – Non-Metals and their Compounds

(vi) Oxygen (O₂)

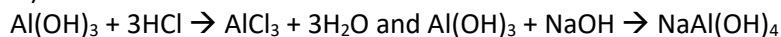
- Oxygen has a simple molecular structure; in its most common form it exists as a diatomic gas (O₂) and makes up 20% of the atmosphere in this form; it also exists as ozone (O₃), especially in the higher parts of the atmosphere (the stratosphere)
- Oxygen is prepared in the laboratory by the decomposition of hydrogen peroxide (see above); it is prepared industrially by the fractional distillation of liquid air (the main components of air are N₂ (79%), O₂ (20%) and Ar (1%); the apparatus is similar to the fractional distillation of crude oil (see Unit 7 – Introduction to Organic Chemistry) but initial temperature used is much lower (-200 °C)
 - the air is cooled to a very low temperature to ensure that all three gases condense into a liquid
 - the liquid air is then run into the base of a fractionating column and the temperature is allowed to rise slowly
 - after argon, nitrogen is the next to boil (-196 °C, 77 K) and it is collected at the top of the fractionating column
 - oxygen-rich liquid is left at the bottom of the fractionating column
- Oxygen is necessary for all life as it is one of the reactants needed in respiration:
$$\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O}$$
This exothermic reaction is the source of energy all plants and animals need to live; oxygen can also react with other fuels, mostly hydrocarbons or alcohols; these are exothermic reactions which produce energy

(vii) Oxides

- The chemical properties of oxides depend on whether the bonding is ionic or covalent; most metals form ionic oxides; all non-metals form covalent oxides; some metals form oxides which are intermediate in character between ionic and covalent
- Metal oxides contain the oxide ion, O²⁻; this is a base as it can accept protons from water to form OH⁻ ions: O²⁻ + H₂O → 2OH⁻; it can also react with acids to form water: O²⁻ + 2H⁺ → H₂O (see Unit 5 – Acids, Bases and Salts); oxides of metals are therefore bases; CaO and CuO are both examples of basic oxides; CaO reacts with water to give the slightly soluble Ca(OH)₂; CuO is insoluble in water but reacts with acids to make salts (see Unit 5 – Acids, Bases and Salts):
$$\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2; \text{CuO} + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{H}_2\text{O}$$
- Non-metal oxides do not contain ions and many react with OH⁻ ions to form salts, or react with water to form H⁺ ions; most oxides of non-metals are therefore acids; examples are: CO₂, SO₂ and SO₃
 - CO₂: dissolves in water to give a slightly acidic solution $\text{CO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$; it reacts with hydroxides to form carbonates: $\text{CO}_2 + 2\text{OH}^- \rightarrow \text{CO}_3^{2-}$; when bubbled through limewater (Ca(OH)₂) it forms CaCO₃ and this is the basis of the test for CO₂ (see Unit 5 – Acids, Bases and Salts)
 - SO₂: dissolves in water to give an acidic solution $\text{SO}_2 + \text{H}_2\text{O} \rightleftharpoons 2\text{H}^+ + \text{SO}_3^{2-}$; SO₂ is sufficiently acidic to turn damp blue litmus paper red and this is the basis of the one of the tests for SO₂ (see Unit 5 – Acids, Bases and Salts)
 - SO₃: dissolves in water to give a strongly acidic solution: $\text{SO}_3 + \text{H}_2\text{O} \rightleftharpoons 2\text{H}^+ + \text{SO}_4^{2-}$; this is how sulphuric acid is made
- The substances formed when covalent oxides react with water are known as **oxo-acids**; they can be considered as hydroxides of non-metals: eg H₂SO₄, HNO₃, HClO

Unit 12 – Non-Metals and their Compounds

- There are a small number of oxides of metals which are amphoteric; they can form salts by reacting with both acids and alkalis; examples are PbO, Al₂O₃ and ZnO (see Unit 5 – Acids, Bases and Salts); metals which form amphoteric oxides also form amphoteric hydroxides; the metals ions form precipitates with OH⁻ ions which dissolve in excess NaOH and this is used as a test for the metal ions (see Unit 8 – Solubility and Precipitation Reactions)



- A small number of oxides are neutral and do not react with acids or bases and therefore cannot be used to make salts; the most common example is carbon monoxide (CO)



Test your knowledge 8.2 – Oxygen and Oxides

- Outline the main stages in the industrial preparation of oxygen
- Write an equation for the reaction of glucose with oxygen and hence explain why oxygen is essential for life
- Explain how you can predict whether an oxide or hydroxide will be acidic or basic
- Classify the following oxides and hydroxides as acidic, basic or amphoteric: MgO, NaOH, CO₂, SO₃, Al₂O₃, ClOH, HPO(OH)₂, FeO, SiO₂
- Write equations for the reactions of CaO with SO₂, NaOH with SiO₂ and CO₂ with MgO

Lesson 9 – What do I need to know about nitrogen and ammonia?

e) Nitrogen and its Compounds



Summary Activity 9.1: What can you remember about nitrogen and ammonia?

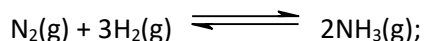
- Describe the structure and bonding in nitrogen
- Describe the structure and bonding in ammonia
- Write an equation to show ammonia behaving as a base
- Describe a simple chemical test for ammonia
- Describe a simple chemical test for ammonium ions
-

(i) Nitrogen (N₂)

- Nitrogen has a simple molecular structure (see Unit 2 – Particles, Bonding and Structure); it exists as a diatomic gas (N₂) and makes up 80% of the atmosphere; the bond between the nitrogen atoms is a triple covalent bond which is very strong and hard to break; this makes nitrogen a very unreactive gas; it is non-polar and so not very soluble in water
- Industrially, pure nitrogen is prepared by the fractional distillation of liquid air in the same process as that used to prepare oxygen industrially (see above)

Unit 12 – Non-Metals and their Compounds

- The main use of nitrogen is to manufacture **ammonia (NH₃)** by the **Haber process**:



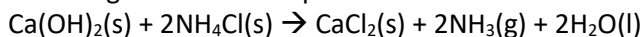
- this is an exothermic and reversible reaction; according to Le Chatelier's principle, the best yield of ammonia will be obtained at high pressure (because the reaction involves a reduction in the number of gas particles) and low temperature (because the reaction is exothermic) (see Unit 4 – Introduction to Physical Chemistry)
- a high pressure of 250 atmospheres (25 MPa) is used
- a very low temperature will cause the reaction to be very slow (see Unit 4 – Introduction to Physical Chemistry); in order to balance the need for a high temperature (to create a fast reaction) and a low temperature (to create a high yield), a compromise temperature of 450 °C is used
- an **iron** catalyst is used in order to speed up the reaction and allow a lower temperature to be used
- the unreacted nitrogen and hydrogen can be recycled

(ii) Ammonia (NH₃) and ammonium ions (NH₄⁺)

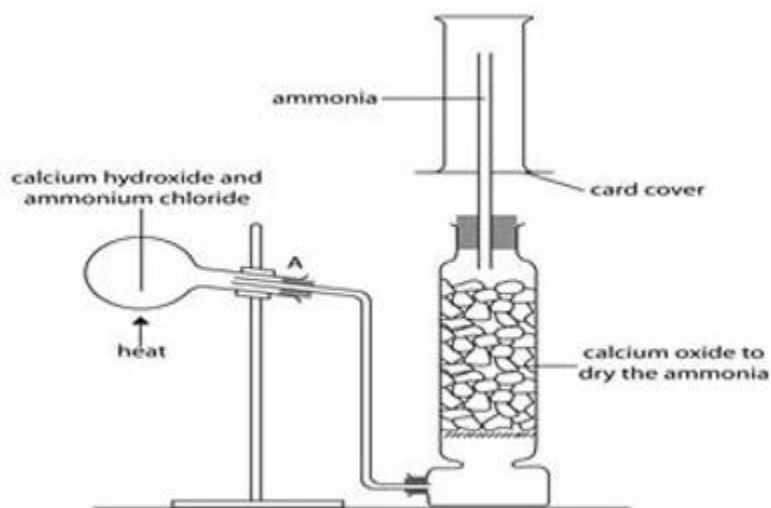
- Ammonia (NH₃)** has a simple molecular structure; it exists as NH₃ molecules and is a pungent-smelling gas; the molecule is pyramidal in shape, is very polar and is able to form hydrogen bonds; as a result it has a higher boiling point (-33 °C or 240 K) than most other small molecules and is very soluble in water
- It reacts with water to form an alkaline solution; $\text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$; it is the only common gas with basic properties; good tests for ammonia gas include its ability to turn red litmus paper blue (because it releases OH⁻ ions in the presence of water) and its ability to neutralise HCl gas, forming a white smoke which is the salt NH₄Cl (see Topic 5 – Acids, Bases and Salts)
- Ammonia can also react with nitric acid and sulphuric acid to make the salts NH₄NO₃ and (NH₄)₂SO₄, which are both important fertilisers; this is the most important use of ammonia
 $\text{NH}_3 + \text{HNO}_3 \rightarrow \text{NH}_4\text{NO}_3$; $2\text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4$
NH₄⁺ ions are a very important source of nitrogen, which plants need; bacteria in the soil gradually oxidise NH₄⁺ ions to NO₃⁻ ions, which plants can absorb through their roots
- Ammonia is also a component of cleaners; alkaline solutions are useful for breaking down fats into water soluble products

Unit 12 – Non-Metals and their Compounds

- The **ammonium ion (NH_4^+)** is acidic and ammonium salts (as salts of weak bases) are therefore acidic; NH_4^+ ions can react with bases to form NH_3 gas; this is used as a test for NH_4^+ ions in solution; if warmed with NaOH solution a pungent gas will be given off which turns damp red litmus paper blue (see Unit 5 – Acids, Bases and Salts); this reaction is also used in the laboratory preparation of NH_3 ; solid $\text{Ca}(\text{OH})_2$ is heated with solid NH_4Cl and the following reaction takes place:



The NH_3 produced also contains water vapour and should be passed through a drying agent such as CaO; it cannot be collected over water as it is very soluble in water; it is collected by upward delivery as it is less dense than air



www.kullabs.com



Demonstration 9.2: Prepare NH_3 gas in the laboratory

Set up the apparatus as shown in the diagram above (a test tube or boiling tube with a bung attached to a delivery tube, a CaO tower and a gas jar, and a Bunsen burner)

- Add two spatulas each of ammonium chloride and calcium hydroxide to the boiling tube; shake the mixture so it is well mixed
- Heat the solid mixture gently using a Bunsen burner until a gas starts to be produced; this gas will be forced through the CaO and be collected in the gas jar, which should be held upside down
- Test the gas by placing damp blue litmus paper into the jar

Write an equation for the reaction taking place

What type of reaction is taking place?

Why can NH_3 not be collected over water?

What is the function of the CaO?

Why must NH_3 be collected by upward delivery?

What is the test for NH_3 and how does it work?

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=95HaDtR35eI

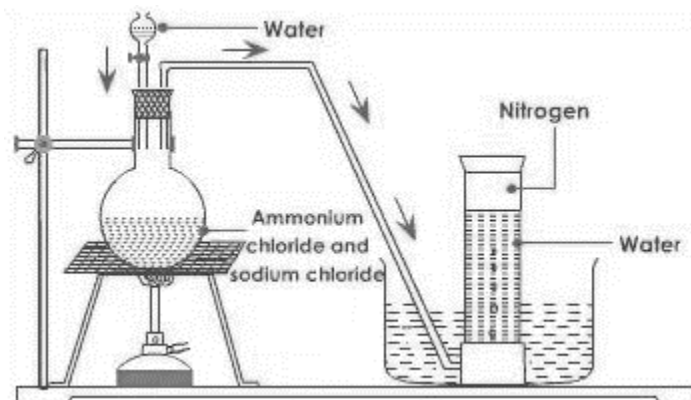
Unit 12 – Non-Metals and their Compounds

- In NH_3 and NH_4^+ , the N has an oxidation number of -3; although N is stable in this form, it can be oxidised to nitrogen gas (oxidation number = 0) by strong oxidising agents; NH_4^+ ions are thus mild reducing agents; one oxidising agent which can be used for this purpose is NaNO_2 ; NO_2^- ions are known as nitrate (III) ions:



The N in NH_4^+ is oxidised from -3 to 0; the N in NO_2^- is reduced from +3 to 0; N_2 gas is thus the product of the oxidation and reduction processes

This is the standard laboratory method for the preparation of nitrogen gas



www.kullabs.com

- There is no good single test for nitrogen gas; a gas is often assumed to be nitrogen if the tests for all other gases are negative; nitrogen puts out a lit splint, has no effect on litmus paper or on the pH of water, and does not readily take part in redox reactions



Demonstration 9.3: Prepare N_2 gas in the laboratory

Set up the apparatus as shown in the diagram above (a round-bottomed flask with a bung attached to a delivery tube, a gas jar, a water trough and a Bunsen burner with tripod and gauze)

- 1) Add around 25 cm^3 of water to the flask
- 2) Add two spatulas each of ammonium chloride and sodium nitrate (III) to the flask
- 3) Heat the mixture gently using a Bunsen burner until a gas starts to be produced; this gas should bubble upwards through the water and be collected in the gas jar

Write an equation for the reaction taking place

What is oxidised and what is reduced in this reaction?

Why can N_2 be collected over water?

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=vpbYWLE9I_Q



Test your knowledge 9.4 – Nitrogen and Ammonia

- Outline the main stages in the industrial preparation of nitrogen
- Describe the laboratory preparation of nitrogen; state the type of reaction taking place and explain why the gas can be collected in this way
- Describe the laboratory preparation of ammonia; state the type of reaction taking place and explain why the gas needs to be collected in this way
- Describe and justify the conditions used in the manufacture of ammonia by the Haber process
- State the main use of ammonia

Lesson 10 – What do I need to know about nitric acid and nitrates?



Summary Activity 10.1: What can you remember about nitric acid and nitrates?

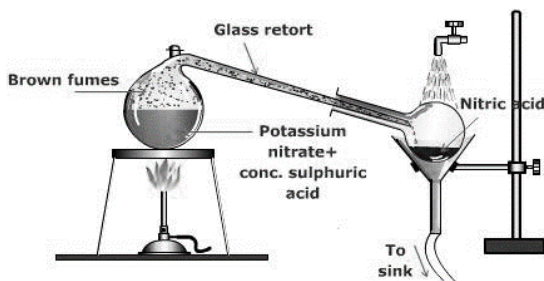
- Give the formula of nitric acid and the nitrate ion
- Discuss the acid-base properties of nitric acid and nitrate ions
- Write an equation to show how nitric acid reacts with ammonia
- Describe a simple chemical test for nitrate ions

(iii) Nitric acid (HNO₃) and nitrate (V) ions (NO₃⁻)

- **Hydrogen nitrate (HNO₃)** has a simple molecular structure; it is a polar molecule with hydrogen bonding and therefore has a relatively high melting point of 83 °C; it is most commonly found dissolved in water; it is highly soluble in water and the mixture is known as **nitric acid**; it is a strong acid, fully dissociated into its ions: H⁺(aq) + NO₃⁻(aq); even concentrated nitric acid contains around 32% water (see Unit 5 – Acids, Bases and Salts)
- Nitric acid reacts with bases to make salts (see Unit 5 – Acids, Bases and Salts); nitrate salts are useful as fertilisers; the NO₃⁻ ion is readily absorbed by plants as a valuable source of nitrogen; this is the most important use of nitric acid; the most commonly used fertiliser is ammonium nitrate, made by reacting nitric acid with ammonia: HNO₃ + NH₃ → NH₄NO₃

Unit 12 – Non-Metals and their Compounds

- Concentrated acids react with salts to make another acid and another salt; the H^+ in the acid is replaced by the cation in the salt and vice versa; this reaction can be used to make a number of salts (see Unit 5 - Acids, Bases and Salts) as well as hydrogen chloride; this reaction is also used in the laboratory preparation of nitric acid:



An animation of this process can be seen on www.youtube.com/watch?v=OxZZA_BbmNs

- Gentle heating of the mixture causes the HNO_3 to evaporate; it is distilled off and can be condensed separately
- Nitrate (V) ions (NO_3^-)** are found in dilute nitric acid and in all salts of nitric acid; in both nitric acid and nitrate (V) ions, the oxidation number of the N is +5; nitrate (V) ions should be distinguished from nitrate (III) ions (NO_2^-) in which the N has an oxidation number of +3; however NO_3^- ions are commonly known simply as “nitrate ions” because they are much more common and stable than NO_2^- ions, which are always referred to as nitrate (III) ions
- All nitrates are soluble in water so do not give precipitation reactions (see Unit 8 – Solubility and Precipitation Reactions); they are salts of a strong acid so are not bases (see Unit 5 – Acids, Bases and Salts); nitrate ions can behave as oxidising agents; when heated with Al and NaOH the NO_3^- ion can be reduced to NH_3 , which is given off as a pungent gas which turns damp red litmus paper blue; the N is reduced from +5 (in the NO_3^-) to -3 (in the NH_3); this is the standard test for NO_3^- ions (see Unit 6 – Redox Reactions)
- Nitrates are unstable on heating – they undergo **thermal decomposition**; the N is reduced and the O is oxidised:
 - NaNO_3 and KNO_3 decompose to give nitrate (III) salts and oxygen gas:
$$2\text{NaNO}_3(\text{s}) \rightarrow 2\text{NaNO}_2(\text{s}) + \text{O}_2(\text{g})$$

The N is reduced from + 5 to +3 and the O is oxidised from -2 to 0
 - All other nitrates decompose to give oxides, $\text{NO}_2(\text{g})$ and $\text{O}_2(\text{g})$:
$$2\text{Ca}(\text{NO}_3)_2(\text{s}) \rightarrow 2\text{CaO}(\text{s}) + 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$$

The N is reduced from +5 to +4 and the O is oxidised from -2 to 0; the NO_2 is clearly visible as a brown gas



Practical 10.2: Decompose different nitrate salts

You are given access to sodium nitrate and calcium nitrate – both salts decompose on heating, but in different ways:

- 1) Place two spatulas of sodium nitrate into a test tube
- 2) Heat the solid sample gently using a Bunsen Burner (medium flame); continue heating until bubbles are seen; hold the test tube at a 45° angle at all times
- 3) Dip a glowing splint into the test tube
- 4) Repeat steps 1 – 3 using calcium nitrate

Write an equation for the reactions occurring when both salts are heated; in each case, state what is oxidised and what is reduced

What happens to the glowing splint in both reactions and why?

What is responsible for the brown colour in the second reaction?

Can't do this experiment? Watch these videos instead: www.youtube.com/watch?v=VNTkJWkPbiM;
www.youtube.com/watch?v=L7qRoYQEBPk



Test your knowledge 10.3 – Nitric acid and nitrates

- (a) Describe how concentrated nitric acid is prepared in the laboratory; name the type of reaction taking place and write an equation for the reaction
- (b) State the main use of nitric acid
- (c) Write an equation to show the reactions taking place when sodium nitrate and copper (II) nitrate are heated separately

Lesson 11 – What do I need to know about sulphur and sulphur dioxide?

f) Sulphur and its Compounds



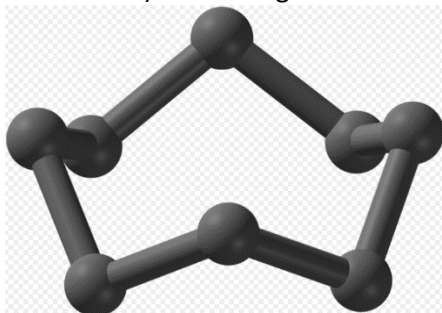
Summary Activity 11.1: What can you remember about sulphur dioxide?

- Give the formula of sulphur dioxide and discuss its acid-base properties
- Describe a simple chemical test for sulphur dioxide

Unit 12 – Non-Metals and their Compounds

(i) Sulphur

- The element sulphur has a simple molecular structure; there are several allotropes of sulphur but most of these contain S₈ molecules, consisting of eight sulphur atoms in a ring; the large size of these molecules creates strong Van der Waal's forces and sulphur is a solid at room temperature with a melting point of 200 °C (473 K); it can also exist as a polymer (with long chains of atoms) and as rings of different sizes



www.wikipedia.com

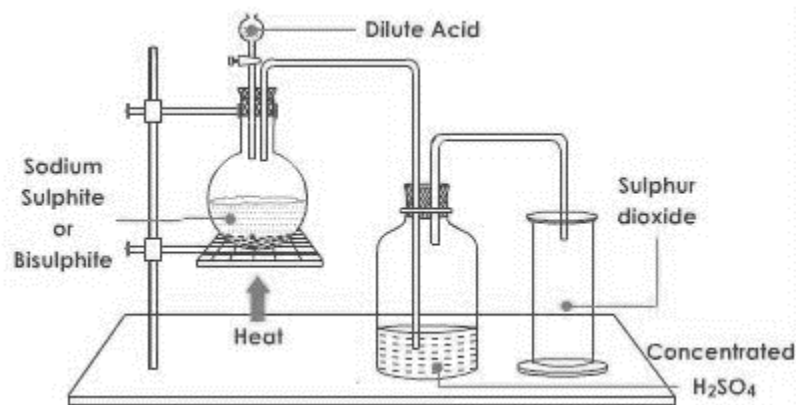
- Sulphur is used in the manufacture of rubber
- Most sulphur is used to manufacture sulphuric acid; the first stage in this process is the reaction of sulphur with oxygen to make sulphur dioxide: $S(s) + O_2(g) \rightarrow SO_2(g)$; many fossil fuels, such as coal and oil, contain sulphur and when the fuel is burned, the sulphur is also burned to make SO₂
- Sulphur also reacts with metals to make **sulphides**, which contain the S²⁻ ion:
 $Zn + S \rightarrow ZnS$
 $Cu + S \rightarrow CuS$

(ii) Sulphur dioxide, sulphuric (IV) acid and sulphate (IV) ions

- Sulphur dioxide (SO₂)** is a colourless gas with a simple molecular structure; like most oxides of non-metals, it is acidic and it will turn damp blue litmus paper red; it dissolves in water to make the weak acid **H₂SO₃ (sulphuric (IV) acid)**:
 $SO_2(g) + H_2O(l) \rightleftharpoons H_2SO_3(aq) \rightleftharpoons 2H^+(aq) + SO_3^{2-}(aq)$
 - This can be used as a test for sulphur dioxide although other acidic gases, such as HCl, also do this (see Unit 5 – Acids, Bases and Salts)
 - SO₂ also reacts with alkalis to make salts containing the **sulphate (IV) ion (SO₃²⁻)**:
 $SO_2 + 2OH^- \rightarrow SO_3^{2-}$

Unit 12 – Non-Metals and their Compounds

- **Sulphate (IV) ions (SO_3^{2-})** react with acids to give off sulphur dioxide gas
 - this can be used a test for sulphate (IV) ions (see Unit 5 – acids, bases and salts); $\text{SO}_3^{2-} + 2\text{H}^+ \rightarrow \text{SO}_2 + \text{H}_2\text{O}$
 - this reaction can also be used in the laboratory preparation of sulphur dioxide; if dilute hydrochloric acid is added to solid sodium sulphate (IV), SO_2 is given off: $\text{Na}_2\text{SO}_3(\text{s}) + \text{HCl}(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$



www.youngstudent16.blogspot.com

- the sulphur dioxide is passed through concentrated H_2SO_4 in order to remove any water vapour; it is collected by downward delivery as it is more dense than air; it cannot be collected over water as it is soluble in water



Demonstration 11.2: Prepare SO_2 gas in the laboratory (Part 1)

(caution – concentrated H_2SO_4 is very corrosive – wear gloves)

- 1) Set up the apparatus as shown in the diagram above (a round-bottomed flask, a dropping funnel, two delivery tubes, two bungs to fit the flasks, each with two holes, a gas jar)
- 2) Add around 20 cm^3 of 1.0 mol dm^{-3} HCl into the dropping funnel; add around 5 g of Na_2SO_3 into the round-bottomed flask; add around 50 cm^3 of concentrated sulphuric acid to the second conical flask
- 3) Gradually add the dilute HCl, and start heating the mixture; you will observe a gas being given off which will eventually collect in the gas jar
- 4) When a reasonable quantity of gas has been collected in the gas jar, close the tap of the dropping funnel, remove the delivery tube from the gas jar and place a lid on the gas jar
- 5) Confirm that the gas is SO_2 by dipping briefly some damp blue litmus paper, and then some acidified dichromate paper, into the gas jar

Write an equation for the reaction taking place

What type of reaction is this?

Why is the gas bubbled through concentrated H_2SO_4 ?

Why is the gas collected by downward delivery?

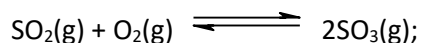
Why is the gas not collected over water?

State and explain the effect of SO_2 on damp blue litmus paper and on acidified dichromate paper

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=x7GycA9_wKg

Unit 12 – Non-Metals and their Compounds

- Sulphate (IV) ions (SO_3^{2-}), sulphuric (IV) acid and sulphur dioxide (SO_2) all contain oxygen in the +4 oxidation state; sulphur is more stable in the +6 oxidation state (in the form of SO_3 , sulphate ions (SO_4^{2-}) and H_2SO_4); all sulphur compounds in the +4 oxidation state are readily oxidised to the +6 oxidation state, and are therefore reducing agents; this is the reason that SO_2 turns dichromate paper from orange to green (see Unit 6 – Redox Reactions)
- The reducing properties of sulphur dioxide is the reason it is added to wine (it prevents the oxidation of the ethanol); it is also the reason that sulphate (IV) salts are added to food (they are antioxidants)
- SO_2 can be oxidised to SO_3 in a very important industrial process called the **Contact Process**:



- this is an exothermic and reversible reaction; according to Le Chatelier's principle, the best yield of SO_3 will be obtained at high pressure (because the reaction involves a reduction in the number of gas particles) and low temperature (because the reaction is exothermic) (see Unit 4 – Introduction to Physical Chemistry)
- in fact a relatively low pressure of 2 atmospheres (200 kPa) is used, because a reasonably good yield is obtained at this pressure
- a very low temperature will cause the reaction to be very slow (see Unit 4 – Introduction to Physical Chemistry); in order to balance the need for a high temperature (to create a fast reaction) and a low temperature (to create a high yield), a compromise temperature of 450 °C is used
- a V_2O_5 catalyst is used in order to speed up the reaction and allow a lower temperature to be used



Test your knowledge 11.3 – sulphur dioxide and sulphate (IV) ions

- Describe a laboratory preparation of sulphur dioxide; write an equation for the reaction and state the type of reaction occurring; explain why the gas is collected by downward delivery and not collected over water
- Explain why SO_2 and SO_3^{2-} ions are reducing agents and state a use for these species which relies on this property
- State and justify the conditions used in the conversion of SO_2 to SO_3 in the Contact process

Lesson 12 – What do I need to know about sulphuric acid and sulphides?



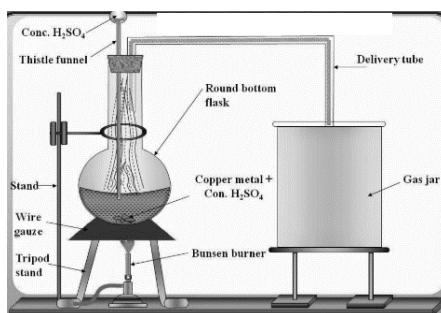
Summary Activity 12.1: What can you remember about sulphuric acid and sulphides?

- Give the formula of sulphuric acid and write an equation to show its reaction with:
 - sodium hydroxide
 - ammonia
 - sodium chloride
 - sodium bromide
 - sodium iodide
- State the type of reaction taking place in each of the above reactions
- Give the formula of sodium sulphide and the sulphide ion
- Describe a simple test for sulphate ions

Unit 12 – Non-Metals and their Compounds

(iii) Sulphur trioxide, Sulphuric acid and Sulphate (VI) ions

- **Sulphur trioxide (SO₃)** is a colourless gas with a simple molecular structure
 - like most oxides of non-metals, it is acidic; it dissolves readily and vigorously in water to make **sulphuric acid (H₂SO₄)**: $\text{SO}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_4(\text{aq})$
 - SO₃ is formed in the atmosphere when SO₂ in the air reacts with oxygen; acid rain is formed when SO₃ in the atmosphere dissolves in rainwater
- **Sulphuric acid (H₂SO₄)** is the acid produced when SO₃ is dissolved in water; it is produced on a huge scale industrially:
 - Sulphur is oxidised to sulphur dioxide ($\text{S} + \text{O}_2 \rightarrow \text{SO}_2$); Sulphur dioxide is oxidised to sulphur trioxide in the Contact Process: $\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$; Sulphur trioxide is dissolved (indirectly) in water to make sulphuric acid
 - Sulphuric acid is a strong acid; it reacts with bases to make **sulphate salts (SO₄²⁻)** (see Topic 5 – Acids, Bases and Salts); the reaction with ammonia is carried out on an industrial scale to make ammonium sulphate, which is an important fertiliser (as a source of nitrogen for plants):
 $2\text{NH}_3(\text{g}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow (\text{NH}_4)_2\text{SO}_4(\text{aq})$
 - Concentrated sulphuric acid also reacts with some salts to make another acid and another salt (see Unit 5 – Acids, Bases and Salts); this is used in the laboratory preparations of HCl and HNO₃ and in the industrial preparation of phosphoric acid:
 $\text{H}_2\text{SO}_4 + 2\text{KCl} \rightarrow \text{K}_2\text{SO}_4 + 2\text{HCl}$
 $\text{H}_2\text{SO}_4 + 2\text{KNO}_3 \rightarrow \text{K}_2\text{SO}_4 + 2\text{HNO}_3$
 $3\text{H}_2\text{SO}_4 + \text{Ca}_3(\text{PO}_4)_2 \rightarrow 3\text{CaSO}_4 + 2\text{H}_3\text{PO}_4$
 - Concentrated sulphuric acid is also an oxidising agent; it can oxidise bromide ions to bromine and iodide ions to iodine:
 $\text{H}_2\text{SO}_4 + 2\text{Br}^- + 2\text{H}^+ \rightarrow \text{SO}_2 + \text{Br}_2 + 2\text{H}_2\text{O}$ (S reduced from +6 to +4; Br oxidised from -1 to 0)
 $\text{H}_2\text{SO}_4 + 8\text{I}^- + 8\text{H}^+ \rightarrow \text{H}_2\text{S} + 4\text{I}_2 + 4\text{H}_2\text{O}$ (S reduced from +6 to -2; Br oxidised from -1 to 0)
Concentrated sulphuric acid can also oxidise metals which are not reactive enough to react with the H⁺ in dilute acids, such as Cu:
 $\text{Cu} + 2\text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{SO}_2 + 2\text{H}_2\text{O}$ (S reduced from +6 to +4; Cu oxidised from 0 to +2)
Note that in most metal acid reactions the H in the acid is reduced (from +1 to 0) (see Unit 6 – Redox Reactions) but in this case the S is reduced; this reaction can be used to prepare sulphur dioxide in the laboratory:



www.youtube.com

- Concentrated sulphuric acid has a very strong affinity for water and can be used as a drying agent, such as in the preparations of chlorine and sulphur dioxide
- Sulphuric acid is also used in car batteries (see Unit 6 – Redox Reactions)



Demonstration 12.2: Prepare SO₂ gas in the laboratory (Part 2)

(caution – concentrated H₂SO₄ is very corrosive; wear gloves)

- 1) Set up the apparatus as shown in the diagram above (a round-bottomed flask, a conical flask, a dropping funnel, a delivery tube, a bung to fit the flask, with two holes, a gas jar, a Bunsen burner with tripod and gauze)
- 2) Add around 20 cm³ of concentrated sulphuric acid into the dropping funnel; add around 5 g of copper turnings into the round-bottomed flask; the concentrated sulphuric acid should just cover the copper turnings (if necessary, add a little more concentrated sulphuric acid)
- 3) Start heating the mixture; you will observe a gas being given off which will collect in the gas jar
- 4) Confirm that the gas is SO₂ by dipping briefly some damp blue litmus paper, and then some acidified dichromate paper, into the gas jar

**Write an equation for the reaction taking place
What is oxidised and what is reduced?**

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=au04TpZ0ON0 (go to 4:15)

- **Sulphate (VI) ions (SO₄²⁻)** are formed when sulphuric acid reacts with bases; they are also formed when SO₂ or SO₃ ions are oxidised
 - the presence of SO₄²⁻ ions in solution can be easily detected by adding dilute hydrochloric acid followed by aqueous barium chloride; a white precipitate of barium sulphate is formed (see Unit 8 – Solubility and Precipitation Reactions): $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$
- sulphur trioxide, sulphuric (VI) acid and sulphate (VI) ions all contain S in the +6 oxidation state; concentrated sulphuric acid is a good oxidising agent but in the absence of acidic conditions S is stable in the +6 oxidation state and sulphate (VI) ions are not good oxidising agents

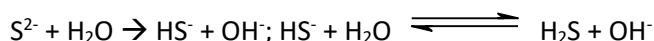


Test your knowledge 12.3 – sulphur trioxide, sulphuric acid and sulphate (VI) ions

- (a) State the acid-base properties of sulphur trioxide and explain how acid rain is produced
- (b) Give a use for sulphuric acid (i) as an acid; (ii) as an oxidising agent; (iii) as a dehydrating agent

(iv) Sulphides (S²⁻)

- Sulphides are formed when sulphur reacts with metals; sulphides are commonly found in minerals in the earth's crust; the most abundant are ZnS, CuS and PbS; sulphides contain the sulphide ion S²⁻
- All metal sulphides (like all metal oxides) are insoluble in water except for Group I sulphides such as Na₂S and K₂S and ammonium sulphide ((NH₄)₂S); the sulphide ion is strongly basic and reacts with water:



Hydrogen sulphide gas (H₂S) is gradually evolved; H₂S smells of rotten eggs; H₂S is also formed when sulphide ions react with acids: eg $\text{CuS}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CuCl}_2(\text{aq}) + \text{H}_2\text{S}(\text{g})$

Unit 12 – Non-Metals and their Compounds

- Metal sulphides and hydrogen sulphide contain S in the -2 oxidation state; sulphides can be oxidised to SO₂ when they are heated in air: eg $2\text{CuS} + 3\text{O}_2 \rightarrow 2\text{CuO} + 2\text{SO}_2$
In these reactions, S is oxidised from -2 to +4 and O is reduced from 0 to -2
This reaction is used to extract metals such as Cu, Pb and Zn from their sulphide ores; it is much easier to extract a metal from its oxide than from its sulphide, so metal sulphides are heated in oxygen to convert them to metal oxides



Test your knowledge 12.4 – sulphide ions and hydrogen sulphide

- Write an equation to show how zinc reacts with sulphur
- Write an equation to show how zinc sulphide reacts with dilute hydrochloric acid
- Write an equation to show how zinc sulphide reacts with oxygen and explain why this reaction is useful
- State how the oxidation number of S changes during the reactions above
- State how hydrogen sulphide gas can be identified

Lesson 13 – What do I need to know about carbon and its oxides?

g) Carbon and its Compounds



Summary Activity 13.1: What can you remember about carbon, carbon dioxide and carbonate ions?

- Describe the structures of diamond and graphite; explain how their different structures result in diamond and graphite having different properties
- Describe how carbon dioxide reacts with water and alkalis
- Describe a test for carbon dioxide gas
- Describe a test for carbonate ions in solution

(i) Allotropes of carbon

- The element carbon exists in a number of forms, called allotropes (allotropes are different structures of the same element); the two most important allotropes of pure carbon are diamond and graphite; the structures of diamond and graphite were covered in Unit 2 (Particles, Bonding and Structures); carbon also exists in other, less pure forms, most notably **coal** and **coke**
- **Coal** is an impure form of carbon; it also contains variable quantities of hydrogen, nitrogen, oxygen and sulphur; coal is a rock formed from dead plant matter which is compressed over time; various forms of coal exist, depending on degree of compression and the time:
 - the plant matter is first converted into **peat** (vegetation which has been partially decayed in acidic, anaerobic conditions; peat contains less than 60% carbon)
 - as the pressure builds up, peat is converted into a sedimentary rock called **lignite** (60 – 70% carbon), which is then converted into **sub-bituminous coal** (71 – 77% carbon) and **bituminous coal** (77 – 87% carbon)
 - the coal eventually being transformed into a metamorphic rock called **anthracite** (more than 87% carbon)
 - these processes take place over millions of years

Unit 12 – Non-Metals and their Compounds

- Peat and coal are burned and used as an energy source ($C + O_2 \rightarrow CO_2$)
- Coal is also converted into other, more useful products by a process called destructive distillation (strong heating in the absence of air); this produces **coke**, **coal gas** (a mixture of gases such as hydrogen, carbon monoxide and methane, which was a widely used fuel before the discovery of natural gas), **coal tar** (a liquid mixture of hydrocarbons used in road surfacing and in some medicines) and **coal oil** (a fuel similar to kerosene)
- **Coke** is a pure form of carbon made by the destructive distillation of coal
 - coke is widely used as a reducing agent in the extraction of metals
 - coke can be converted into synthesis gas (a mixture of carbon monoxide and hydrogen) by a process called gasefication (heating at $700\text{ }^\circ\text{C}$ in a limited supply of oxygen) ($C + H_2O \rightarrow CO + H_2$); synthesis gas is itself a fuel and can also be used in the manufacture of methanol and hydrogen)

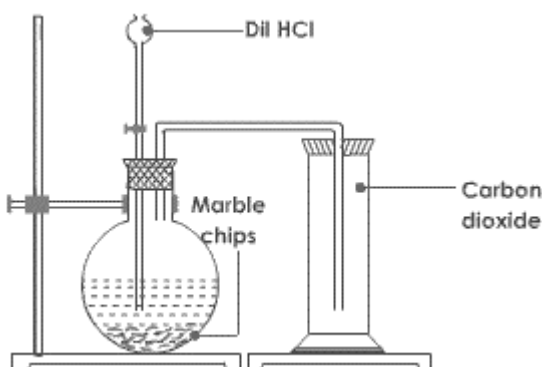


Test your knowledge 13.2 – different forms of carbon

- Explain the difference between peat, lignite, coal and anthracite
- Explain how coke is produced from coal and list any other useful products of this process
- What is synthesis gas and how is it made?

(ii) Carbon dioxide (CO_2), carbonate ions (CO_3^{2-}) and hydrogencarbonate ions (HCO_3^-)

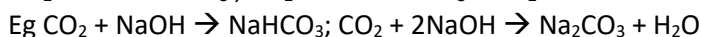
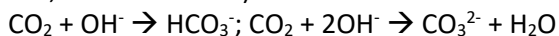
- **Carbon dioxide (CO_2)** is the most stable oxide of carbon; it has a linear simple molecular structure; the C atom is attached to both O atoms with a double covalent bond $O=C=O$
 - Carbon dioxide can be prepared by direct combination of carbon and oxygen ($C + O_2 \rightarrow CO_2$) but is generally prepared by reacting carbonates with acids (see Unit 5 – Acids, Bases and Salts):
 $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2 + H_2O$



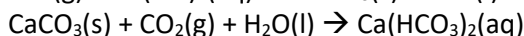
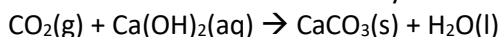
www.igcse2016.wordpress.com

Unit 12 – Non-Metals and their Compounds

- Carbon dioxide is acidic; it reacts with alkalis to form hydrogencarbonate ions and carbonate ions (see Unit 5: acids, bases and salts)



This reaction is the basis for the test for carbon dioxide: CO_2 reacts with limewater ($\text{Ca}(\text{OH})_2$), initially to make calcium carbonate (which is insoluble) and then calcium hydrogencarbonate (which is soluble), so carbon dioxide turns limewater milky and then colourless again (see Unit 5 – Acids, Bases and Salts):



- Carbon dioxide is not flammable at all and is used in fire extinguishers; because it is more dense than air, it sinks onto the fire, preventing oxygen from getting to the fuel

- **Carbonate ions (CO_3^{2-})** are commonly found in rocks

- carbonate ions are formed when metal oxides (which are basic) react with CO_2 (which is acidic)



- carbonate ions can also be formed in solution by dissolving carbon dioxide in alkalis

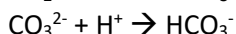
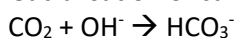


- carbonate ions react with acids to form carbon dioxide: $\text{CO}_3^{2-} + 2\text{H}^+ \rightarrow \text{CO}_2 + \text{H}_2\text{O}$ (see Unit 5 – Acids, Bases and Salts); this is commonly used in the preparation of carbon dioxide

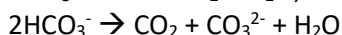
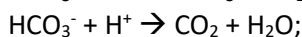
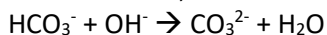
- **Sodium carbonate (Na_2CO_3)** is a weak base, due to the CO_3^{2-} ion; it is also known as washing soda; it is one of a small number of soluble carbonates and so is a useful source of CO_3^{2-} in aqueous solution; it is used in the manufacture of glass (by mixing it with silicon dioxide); it is also used as a water softener and to raise the pH of water; it is also very useful for making standard solutions because it is stable in its most common form $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$

- **Calcium carbonate (CaCO_3)** is a weak base due to the CO_3^{2-} ion; it is insoluble in water and is the main component of limestone, chalk and marble; it is added to soil to reduce its acidity and is widely used in building materials, although it is damaged by acids; it is also used in cement; it decomposes on heating into CaO ; $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$; it is therefore used to make CaO and is generally added during the manufacture of steel to produce the CaO

- **Hydrogencarbonate ions (HCO_3^-)** are bases made by the partial neutralisation of CO_2 with alkalis or by the partial neutralisation of carbonate ions with acids:



- HCO_3^- ions react with bases to make carbonates and with acids to make CO_2 ; they are therefore amphoteric (see Unit 5 – Acids, Bases and Salts); HCO_3^- ions also react with themselves to make carbonates and CO_2 :



- **Sodium hydrogencarbonate (NaHCO_3)** is a weak base due to the HCO_3^- ion; it is also known as baking soda or baking powder; NaHCO_3 reacts with itself on heating to release CO_2 : $2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$; when added to dough, the CO_2 is slowly released into the dough on heating, causing it to rise; it is therefore very useful in baking; in water it behaves predominantly as a mild base and is taken in tablets to reduce stomach acidity (ie treat indigestion); it is also used to raise the pH of swimming pools



Demonstration 13.3: Prepare CO₂ gas in the laboratory

- 1) Set up the apparatus as shown in the diagram above (a round-bottomed flask, a dropping funnel, a delivery tube, a bung to fit the flask, with two holes, a gas jar)
- 2) Add around 50 cm³ of 2 moldm⁻³ HCl into the dropping funnel; add around 5 g of marble chips into the round-bottomed flask
- 3) Whilst the gas is being produced, dip the delivery tube into a test tube containing water for a few minutes and observe what happens
- 4) Collect a small amount of the gas in a test tube. Light a splint and place it into the test tube; observe what happens

Write an equation for the reaction taking place

What type of reaction is this?

Why is the gas collected using downward delivery?

What effect does CO₂ have on limewater and why?

What effect does CO₂ have on a lit splint and why?

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=8zF7nKX6wOY



Test your knowledge 13.4 – carbon dioxide, carbonates and hydrogencarbonates

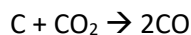
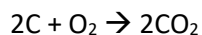
- (a) Write equations to show the acidic behaviour of carbon dioxide, the basic behaviour of carbonate ions and the amphoteric behaviour of hydrogencarbonate ions
- (b) Describe the laboratory preparation of carbon dioxide and explain how the gas is collected
- (c) Explain why carbon dioxide is used in fire extinguishers
- (d) Give the formulae of sodium carbonate, calcium carbonate and sodium hydrogencarbonate and explain why each is useful

(iv) Carbon monoxide

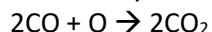
- Carbon monoxide (CO) is a less stable oxide of carbon; it has a simple molecular structure similar to nitrogen, with a triple bond between the C and the O; the third bond is a dative covalent bond, with O providing both

electrons: $\text{C} \equiv \text{O}$

- CO is formed when carbon or hydrocarbons react with a limited supply of oxygen, or by reacting carbon with carbon dioxide:



- it is a colourless and odourless gas and it is extremely toxic, and it binds irreversibly with red blood cells, preventing them from carrying oxygen; it is therefore important that carbon and hydrocarbons are not burned in confined spaces in which people live
- it reacts readily with oxygen to produce CO₂:



Lesson 14 – What do I need to know about rocks?

h) Rocks

- A **mineral** is an element or compound of fixed composition which is found naturally in the earth's crust and not produced by life processes
- A **rock** is a natural substance made from one or more minerals and found in the earth's crust; rocks are composed of grains of minerals
- The most abundant mineral in the earth's crust is silicon dioxide (SiO_2), but many other minerals contain metals in the form of oxides, carbonates or salts; elements occurring naturally are also classified as minerals

(i) Types of rock

- Rocks can be classified into three types based on their texture and how they are formed: igneous, metamorphic and sedimentary
- **Igneous** rocks are formed when molten rock cools and solidifies; molten rock above ground (lava) cools quickly to form rocks with small crystals; these are called extrusive igneous rocks (eg basalt and pumice); molten rock below ground (magma) cools slowly to form rocks with large crystals; these are called intrusive igneous rocks (eg granite)
- **Sedimentary** rocks are formed fragments of other rocks and organisms accumulate and cement together and are then further compacted, either on the earth's surface or under water (eg limestone, shale and sandstone)
- **Metamorphic** rocks are formed when sedimentary rocks are subjected to pressures and temperatures higher than that which resulted in their formation and which result in a significant change in the structure of the rock (eg slate and marble)

(ii) Weathering

- **Weathering** is the breaking down of rocks into smaller particles; the smaller particles are transported to other areas by wind, water and other land movement; the combination of weathering and **transport** is called **erosion**
- There are three types of weathering: physical, biological and chemical:
 - Physical weathering is the fragmentation of rocks as a result of a physical process; the most common is freeze-thaw weathering, which consists of water getting into rock cracks and then freezing; the resulting expansion cracks and fragments the rock; the movement of water against the coast and in rivers, and the movement of rocks against each other, are also causes of physical weathering
 - Biological weathering is the fragmentation of rocks by the action of living organisms; the most important is the growth of tree and other plant roots through rocks, which causes them to fragment
 - Chemical weathering is the fragmentation of rocks as a result of chemical processes:
Hydrolysis is the chemical breakdown of a rock as a result of a reaction with H^+ , OH^- or water; these ions can react with silica rocks, change their chemical composition and weaken their structure

Unit 12 – Non-Metals and their Compounds

Hydration is the absorption of water by a rock to form a new substance; oxides can react with water to form hydroxides and some minerals can absorb water and become hydrated; some silica rocks (especially clays) can absorb water; this can change their composition and weaken their structure

Carbonation is the dissolving of carbon dioxide in rainwater to form carbonic acid; the resulting rainwater is acidic and will dissolve most carbonate rocks to form soluble salts

Oxidation is the reaction of a rock with oxygen in the air; this is common in rocks with contain iron (the Fe^{2+} is oxidised to Fe^{3+}); the resulting rock is usually softer and weaker than the original rock; rusting is a type of oxidation



Test your knowledge 14.1 – rocks

- Explain the difference between a rock and a mineral
- Describe the three types of rock and explain how each is formed
- Explain what is meant by the term “weathering”
- Give three examples of chemical weathering

Unit 12 – Non-Metals and their Compounds

Lesson 15 – What have I learned about non-metals and their compounds?



15.1 END-OF-UNIT QUIZ

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

1. Describe the structure and bonding in argon, chlorine, hydrogen, oxygen, nitrogen and sulphur.
2. Describe the laboratory preparations of oxygen, chlorine and nitrogen; state the type of reaction taking place and explain, with a reason, how the gas is collected in each case.
3. Describe the laboratory preparations of HCl, NH₃ and CO₂; state the type of reaction taking place and explain, with a reason, how the gas is collected in each case.
4. Describe the tests for hydrogen, oxygen, chlorine and nitrogen.
5. Write equations for the following reactions:
 - (a) Chlorine with sodium hydroxide
 - (b) Chlorine with sodium bromide
6. Describe how nitrogen and oxygen are prepared industrially from air.
7. Describe the industrial preparation of ammonia, justifying the conditions used.
8. Write equations to show what happens when the following substances are heated:
 - (a) Calcium nitrate
 - (b) Sodium nitrate
 - (c) Copper sulphide in the presence of air
9. Describe the stages in the formation of anthracite from peat.
10. Name the three types of rock and explain how they are formed.
11. What is weathering? Name three types of chemical weathering.