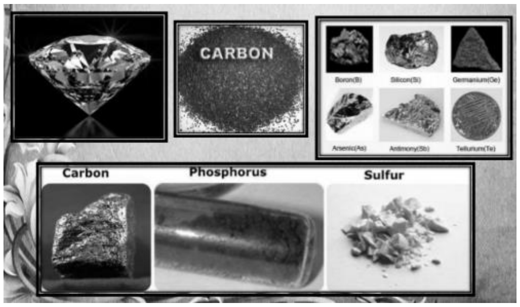
UNIT 12

NON-METALS AND THEIR COMPOUNDS

Student Version



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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: 14 hours

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

UNIT 12 SUMMARY AND SYLLABUS REFERENCE

Lesson	Title and Syllabus Reference
1	**Introduction to Non-Metals; Noble Gases; Physical Properties of Halogens** ***CA2bi periodic gradation of elements in group i.e. the halogens: $F_2/Cl_2/Br_2/I_2$*** *(recognition of group variations noting any anomalies, treatment should include physical states, melting and boiling points);* ***CC7ci halogens - chlorine: properties****;* ***CC7f the noble gases – properties and uses****;* ***ISA4.1 particulate nature of matter*** *(elements - metals and non-metals (1st to 20th elements in the periodic table));* ***ISA10.1 classification of materials*** *(classification of materials into metals, semi-metals (metalloids), and non-metals)* ***ISA10.2 uses of metals and semi-metals*** *(physical properties of non-metals under conductivity, luster, malleability, ductility, sonority, density, melting point and tensile strength)*
2	**Chemical Properties of Halogens and Halides I** ***CA2bi periodic gradation of elements in group i.e. the halogens: $F_2/Cl_2/Br_2/I_2$*** *(recognition of group variations noting any anomalies, treatment should include redox properties of the elements: displacement reaction of one halogen by another);* ***CC7ci halogens - chlorine: properties and reactions****;* ***CC8a properties and uses of sodium and its compounds*** *(NaCl)*
3	**The Reducing Power of Halides** ***CA13ci characteristic reaction of conc. H2SO4 on solid samples of Cl-*** *(no formal scheme of analysis is required);* ***CC7ci halogens - chlorine: laboratory preparation****;* ***CC8a properties and uses of sodium and its compounds*** *(NaCl)*
4	**Chemical Properties of Halogens and Halides III** ***CA2bii properties of chlorine as a typical halogen*** *(properties should include variable oxidation states, reaction with water and alkali (balanced equations required));* ***CA8b practical application of solubility*** *(generalisations of solubility of salts and their applications in qualitative analyses);* ***CC7ci halogens - chlorine: properties and reactions****;* ***CC7cii halogens - uses of halogen compounds*** *(uses should include silver halide in photography and sodium oxochlorate (I) as a bleaching agent);* ***CC8a properties and uses of sodium and its compounds*** *(NaClO)*
5	**Water** ***CC7bii water and solution - water as a solvent****;* ***CC7biv water and solution - treatment of water for town supply*** *(test for water will be required);* ***ISA9.1 physical and chemical properties of water*** *(experiments to determine/demonstrate the boiling point of water, the solvent action of water on a variety of substances, the polar nature of water, the presence of dissolved substances, uses of water);* ***ISA9.3 treatment of water for public consumption*** *(steps involved in the treatment of water for public consumption)*
6	**Hydrogen, Oxygen and their other Compounds** ***CA5bii the gases - properties of gases*** *(chemical properties of H_2);* ***CA7a definitions of acids*** *(effects of acids on metals and indicators);* ***CA7b chemical properties of acids*** *(balanced chemical equations of all reactions);* ***CA13cii characteristic reaction of dilute HCl on solids or aqueous solutions of SO_3^{2-} and CO_3^{2-}*** *(no formal scheme of analysis is required);* ***CA13civ characteristic test for H_2*** *(no formal scheme of analysis is required);* ***CC7ai oxygen – laboratory and industrial preparation****;* ***CC7aii oxygen - properties and uses****;* ***CC7aiii oxygen - binary compounds of oxygen: acidic oxides, basic oxides, amphoteric oxides and neutral oxides****;* ***ISA7.2 chemical properties of acids, bases and salts*** *(test for H_2 gas);* ***ISA10.2 uses of non-metals*** *(uses of O_2);* ***ISE6.2 first aid methods*** *(demonstration of the following using models: methods of extinguishing different fires)*
7	**Sulphur and its Compounds** ***CA5bii the gases – preparation and properties of*** *gases (chemical properties of SO_2);* ***CA8b practical application of solubility*** *(generalisations of solubility of salts and their applications in qualitative analyses);* ***CA13cii characteristic reaction of dilute HCl on solids or aqueous solutions of SO_3^{2-}, confirmatory test for SO_3^{2-} and of conc. H_2SO_4 on solid samples of Cl^-*** *(no formal scheme of analysis is required);* ***CA13civ characteristic test for SO_2 gas*** *(no formal scheme of analysis is required);* ***CC7ei Sulphur – allotropes and uses****;* ***CC7eii sulphur – compounds of sulphur: sulphides****;* ***CC7eii sulphur - compounds of sulphur: trioxosulphate (IV) acid and its salts****;* ***sulphur - compounds of sulphur: tetraoxosulphate (VI) acid - industrial preparation, reactions and uses*** *(for industrial preparation, only the Contact Process should be discussed)*

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Lesson	Title and Syllabus Reference
8	Carbon and Its Inorganic Compounds ***CA5bii the gases - properties of gases*** *(chemical properties of CO_2);* ***CA13cii characteristic reaction of dilute HCl on solids or aqueous solutions of CO_3^{2-}, confirmatory test for CO_3^{2-}*** *(no formal scheme of analysis is required);* ***CA13civ characteristic test for CO_2 gas*** *(no formal scheme of analysis is required);* ***CC6a allotropes of carbon other than diamond and graphite*** *(structures, properties and uses, the uses of the allotropes should be correlated with their properties and structures, combustion of allotropes);* ***CC6bi coal – different types*** *(different types should include anthracite, peat and lignite);* ***CC6bii coal - destructive distillation of coal and uses of the products****;* ***CC6ci coke - gasefication and uses****;* ***CC6cii coke - manufacture of synthetic gas and uses****;* ***CC6di oxides of carbon - carbon (IV) oxide: uses****;* ***CC6dii oxides of carbon - carbon (II) oxide: properties and uses****;* ***CC8a properties and uses of sodium and its compounds*** *(Na_2CO_3 and $NaHCO_3$);* ***CC8b properties and uses of calcium and its compounds*** *($CaCO_3$);* ***ISA7.2 chemical properties of acids, bases and salts*** *(test for carbon dioxide gas);* ***ISA10.2 uses of non-metals*** *(uses of C)*
9	Nitrogen and Ammonia ***CA5c the gases - properties of gases*** *(chemical properties of NH_3);* ***CA13ci characteristic tests of NH_4^+ with NaOH, confirmatory test for NH_4^+*** *(no formal scheme of analysis is required);* ***CA13civ characteristic test for NH_3 gas*** *(no formal scheme of analysis is required);* ***CC7di nitrogen - preparation and properties*** *(both laboratory and industrial preparations from liquefied air are required);* ***CC7dii nitrogen – uses of nitrogen****;* ***CC7diii nitrogen - compounds of nitrogen – ammonia: industrial preparation and uses****;* ***ISA7.2 chemical properties of acids, bases and salts*** *(test for ammonia gas);* ***ISA10.2 uses of non-metals*** *(uses of N_2)*
10	Nitric Acid and Nitrates ***CA13cii characteristic reaction of conc. H_2SO_4 on solid samples of NO_3^-; confirmatory test for NO_3^-*** *(no formal scheme of analysis is required);* ***CC7diii nitrogen - compounds of nitrogen: trioxonitrate (V) acid - laboratory preparation, reactions and uses****;* ***nitrogen - compounds of nitrogen: trioxonitrate (V) salts*** *(action of heat will be required);* ***CC8a properties and uses of sodium and its compounds*** *($NaNO_3$)*
11	The preparation and collection of gases I ***CA5bii the gases - preparation of gases*** *(laboratory preparation of gases lighter than air (H_2) to illustrate the principles of purification and collection of gases;* ***ISA7.2 chemical properties of acids, bases and salts*** *(description of laboratory preparation of hydrogen gas, test for hydrogen gas);* ***CC7ai oxygen - laboratory preparation; CC7ci halogens - chlorine: laboratory preparation****;* ***CC7di nitrogen - preparation*** *(laboratory preparation is required)*
12	The preparation and collection of gases II ***CA5bii the gases - preparation of gases*** *(laboratory preparation of gases lighter than air (NH_3) and heavier than air (HCl, SO_2, CO_2) to illustrate the principles of purification and collection of gases);* ***CA13civ characteristic test for HCl*** *(no formal scheme of analysis is required);* ***ISA7.2 chemical properties of acids, bases and salts*** *(description of laboratory preparation and test for ammonia gas and CO_2 gas)*
13	Rocks ***ISA6.1 types, formation and characteristics of rocks*** *(formation of igneous, sedimentary and metamorphic rocks and their characteristics);* ***ISA6.2 weathering of rocks*** *(physical, biological and chemical weathering of rocks, explanation of the effect of hydration, hydrolysis, carbonation and oxidation on rocks is required)*
14	Unit 12 Revision and Summary

The Periodic Table of the Elements

Key

relative atomic mass
symbol
name
atomic (proton) number

1	2											3	4	5	6	7	0
							1.0 **H** hydrogen 1										(18) 4.0 **He** helium 2
(1)	(2)											(13)	(14)	(15)	(16)	(17)	
6.9 **Li** lithium 3	9.0 **Be** beryllium 4											10.8 **B** boron 5	12.0 **C** carbon 6	14.0 **N** nitrogen 7	16.0 **O** oxygen 8	19.0 **F** fluorine 9	20.2 **Ne** neon 10
23.0 **Na** sodium 11	24.3 **Mg** magnesium 12	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	27.0 **Al** aluminium 13	28.1 **Si** silicon 14	31.0 **P** phosphorus 15	32.1 **S** sulfur 16	35.5 **Cl** chlorine 17	39.9 **Ar** argon 18
39.1 **K** potassium 19	40.1 **Ca** calcium 20	45.0 **Sc** scandium 21	47.9 **Ti** titanium 22	50.9 **V** vanadium 23	52.0 **Cr** chromium 24	54.9 **Mn** manganese 25	55.8 **Fe** iron 26	58.9 **Co** cobalt 27	58.7 **Ni** nickel 28	63.5 **Cu** copper 29	65.4 **Zn** zinc 30	69.7 **Ga** gallium 31	72.6 **Ge** germanium 32	74.9 **As** arsenic 33	79.0 **Se** selenium 34	79.9 **Br** bromine 35	83.8 **Kr** krypton 36
85.5 **Rb** rubidium 37	87.6 **Sr** strontium 38	88.9 **Y** yttrium 39	91.2 **Zr** zirconium 40	92.9 **Nb** niobium 41	96.0 **Mo** molybdenum 42	[98] **Tc** technetium 43	101.1 **Ru** ruthenium 44	102.9 **Rh** rhodium 45	106.4 **Pd** palladium 46	107.9 **Ag** silver 47	112.4 **Cd** cadmium 48	114.8 **In** indium 49	118.7 **Sn** tin 50	121.8 **Sb** antimony 51	127.6 **Te** tellurium 52	126.9 **I** iodine 53	131.3 **Xe** xenon 54
132.9 **Cs** caesium 55	137.3 **Ba** barium 56	138.9 **La** * lanthanum 57	178.5 **Hf** hafnium 72	180.9 **Ta** tantalum 73	183.8 **W** tungsten 74	186.2 **Re** rhenium 75	190.2 **Os** osmium 76	192.2 **Ir** iridium 77	195.1 **Pt** platinum 78	197.0 **Au** gold 79	200.6 **Hg** mercury 80	204.4 **Tl** thallium 81	207.2 **Pb** lead 82	209.0 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac** † actinium 89	[267] **Rf** rutherfordium 104	[268] **Db** dubnium 105	[271] **Sg** seaborgium 106	[272] **Bh** bohrium 107	[270] **Hs** hassium 108	[276] **Mt** meitnerium 109	[281] **Ds** darmstadtium 110	[280] **Rg** roentgenium 111	Elements with atomic numbers 112-116 have been reported but not fully authenticated						

* **58 – 71** Lanthanides	140.1 **Ce** cerium 58	140.9 **Pr** praseodymium 59	144.2 **Nd** neodymium 60	[145] **Pm** promethium 61	150.4 **Sm** samarium 62	152.0 **Eu** europium 63	157.3 **Gd** gadolinium 64	158.9 **Tb** terbium 65	162.5 **Dy** dysprosium 66	164.9 **Ho** holmium 67	167.3 **Er** erbium 68	168.9 **Tm** thulium 69	173.1 **Yb** ytterbium 70	175.0 **Lu** lutetium 71
† **90 – 103** Actinides	232.0 **Th** thorium 90	231.0 **Pa** protactinium 91	238.0 **U** uranium 92	[237] **Np** neptunium 93	[244] **Pu** plutonium 94	[243] **Am** americium 95	[247] **Cm** curium 96	[247] **Bk** berkelium 97	[251] **Cf** californium 98	[252] **Es** einsteinium 99	[257] **Fm** fermium 100	[258] **Md** mendelevium 101	[259] **No** nobelium 102	[262] **Lr** lawrencium 103

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

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 ^{o}C) and Xe boils at 166 K (-107 ^{o}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 denser than air

- 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 – understanding 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

- 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 – understanding the halogens (physical properties)

(a) State the colour of pure iodine, iodine in cyclohexane and iodine in water
(b) State the colours of chlorine and bromine when dissolved in cyclohexane or water
(c) State and explain the trend in boiling points of the halogens
(d) 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):

$$F_2(g) + 2e \rightleftharpoons 2F^-(aq)$$
$$Cl_2(g) + 2e \rightleftharpoons 2Cl^-(aq)$$
$$Br_2(l) + 2e \rightleftharpoons 2Br^-(aq)$$
$$I_2(aq) + 2e \rightleftharpoons 2I^-(aq)$$

- 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

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

$$2F^-(aq) \rightleftharpoons F_2(g) + 2e$$
$$2Cl^-(aq) \rightleftharpoons Cl_2(g) + 2e$$
$$2Br^-(aq) \rightleftharpoons Br_2(aq) + 2e$$
$$2I^-(aq) \rightleftharpoons I_2(aq) + 2e$$

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

Eg $Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$

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

$Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$

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

$Br_2(aq) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(aq)$

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

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

(a) State and explain the trend in oxidising power of the halogens
(b) State and explain the trend in reducing power of the halides
(c) 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
(d) State what you would observe when bromine is added to a solution of sodium chloride; explain your observations

Practical 2.2: Carry out halogen displacement reactions

Note: take care not to inhale the fumes of Cl_2 or Br_2; follow teacher's instructions when disposing of chemicals for this experiment

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 cm^3 of KCl(aq) into two separate test tubes
2) Add 2 cm^3 of bromine water to one of the test tubes and 2 cm^3 of iodine water to the other test tube and record your observations
3) Add 1 cm^3 of cyclohexane to both test tubes and shake well; record your observations
4) Pour 2 cm^3 of KBr(aq) into two separate test tubes
5) Add 2 cm^3 of chlorine water to one of the test tubes and 2 cm^3 of iodine water to the other test tube and record your observations
6) Add 1 cm^3 of cyclohexane to both test tubes and shake well; record your observations
7) Pour 2 cm^3 of KI(aq) into two separate test tubes
8) Add 2 cm^3 of chlorine water to one of the test tubes and 2 cm^3 of bromine water to the other test tube and record your observations
9) Add 1 cm^3 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 – How else can we compare the reducing properties of halide ions?

- Further evidence for the difference in reducing power of the halide ions comes from the different ways in which solid samples of halides 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
    $2H_2SO_4(l) + 2KBr(s) \rightarrow SO_2(g) + Br_2(g) + 2H_2O(l) + K_2SO_4(s)$
    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)
    $5H_2SO_4(l) + 8KI(s) \rightarrow H_2S(g) + 4I_2(g) + 4H_2O(l) + 4K_2SO_4(s)$
    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:
    $H_2SO_4(l) + 2KCl(s) \rightarrow K_2SO_4(s) + 2HCl(g)$
    this is not a redox reaction; white fumes of HCl are seen; **this is the standard laboratory method for the preparation of HCl gas**

- It is possible to oxidise chloride ions to chlorine in the presence of a very strong oxidising agent, such as MnO_2:
  $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$
  - the Mn is reduced from +4 to +2; the Cl^- is oxidised to Cl_2 (-1 to 0)
  - **this is the standard laboratory preparation of Cl_2 gas**

Test your knowledge 3.1: Group 7 – comparing the reducing power of halides

(a) Write an equation for the reaction of potassium chloride with concentrated sulphuric acid
(b) Write an equation for the reaction of potassium bromide with concentrated sulphuric acid
(c) Write an equation for the reaction of potassium iodide with concentrated sulphuric acid
(d) Explain how the equations in (g), (h) and (i) illustrate the trend in reducing properties of the halides
(e) Identify an oxidising agent which can oxidise chloride ions to chlorine. Would you expect this oxidising agent to be able to oxidise bromide and iodide ions?

Demonstration 3.2: Compare the reactions of concentrated sulphuric acid with the different halide ions

(CAUTION – the reaction releases toxic gases – use a fume cupboard; concentrated H_2SO_4 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 cm^3 of concentrated H_2SO_4; note the colour of any gases produced
3) Test the vapours produced with ammonia and 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 else do I need to know about halogens and halides?

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

- Describe a simple test to identify chloride, bromide and iodide ions in solution

- Chlorine will combine directly with most metals to make metal chlorides; the chlorine is reduced (from 0 to -1) and the metal is oxidised:
  Eg $Mg(s) + Cl(g) \rightarrow MgCl_2(s)$; $2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$
  Bromine and iodine can react with some of the more reactive metals but the reactions tend to be slower and less vigorous

- Chlorine is slightly soluble in water and also reacts slightly with water:
  $$Cl_2(g) + H_2O(l) \rightleftharpoons HCl(aq) + HClO(aq)$$
  - 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:
  $$Cl_2(g) + 2NaOH(aq) \rightarrow NaCl(aq) + NaClO(aq) + H_2O(l)$$
  - This reaction is similar to the reaction of chlorine with water; 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))**
  $$ClO^- + 2H^+ + 2e^- \rightarrow Cl^- + H_2O$$

- Chloride, bromide and iodide ions all from precipitates with the silver (I) ion, Ag^+; each has its characteristic colour and can 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
  $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$ white precipitate
  $Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$ cream precipitate
  $Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$ yellow precipitate

(iii) Uses of Halogens and Halides

- **Chlorine** is added to water as a sterilising agent because of its oxidising properties; i**odine**, 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: $ClO^- + 2H+ + 2e^- \rightarrow Cl^- + H_2O$

- **Silver bromide (AgBr)** is used in conventional photography; it is applied to photographic film and decomposes on exposure to light, leaving silver: $2AgBr(s) \rightarrow 2Ag(s) + Br_2(l)$; 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 4.2: Group 7 – describing other reactions of halogens and halides

(a) Magnesium reacts with both chlorine and bromine; write equations for the reactions taking place
(b) Explain why magnesium reacts faster with chlorine than with bromine
(c) Write an equation for the reaction of chlorine with water and hence explain why chlorine is added to swimming pools and drinking water
(d) Write an equation for the reaction of chlorine with dilute sodium hydroxide and explain why the product is useful
(e) Describe how the oxidation number of Cl changes during the reactions in (c) and (d)
(f) Describe a simple laboratory test for chlorine gas
(g) Describe a simple test for chloride, bromide and iodide ions in aqueous solution
(h) Give a use for chlorine, iodine, sodium chlorate (I) and silver bromide

Lesson 5 – Why is water special?

d) Water

Summary Activity 5.1: The special properties of water

- Describe the structure and bonding in water
- Describe how you would measure the boiling point of water and explain why the boiling point of water is unusually high
- Which substances generally dissolve in water and which do not?

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

- 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 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 5.2: Demonstrate the polarity of water

1) Inflate a balloon and rub it against an item of clothing (this makes it electrostatically charged)
2) Turn a tap on and allow a steady stream of water to run
3) 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=8RSDK17lnUc

- 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)

Practical 5.3: Investigate the solubility of different substances in water

1) Place seven test tubes into a test tube rack and place 5 cm^3 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 cm^3 of ethanol and into the seventh, 2 cm^3 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?

- 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

- Societies need to 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

Test your knowledge 5.4: Understanding the Importance of Water

(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

Lesson 6 – What do I need to know about hydrogen, oxygen and their compounds?

e) Hydrogen, Oxygen and their Other Compounds

Summary Activity 5.1: What do you know about hydrogen, oxygen and their compounds?

- Describe the structure and bonding in hydrogen and oxygen
- 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
- Describe a simple test for OH^- ions in aqueous solution
- 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

(i) Hydrogen (H_2)

- 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

- 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

(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 $HCl \rightarrow H^+ + Cl^-$

- H^+ 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)

- H^+ ions react with metals above hydrogen in the reactivity series (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):
  $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$ or $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
  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; **this reaction is used in the laboratory preparation of hydrogen gas**

(iii) Hydroxide ions (OH^-)

- Water reacts with bases to produce solutions containing hydroxide ions: eg $NH_3 + H_2O == NH_4^+ + OH^-$ (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)

(iv) Hydrogen Peroxide (H_2O_2)

- The most common oxide of hydrogen is water (H_2O); this contains oxygen in an oxidation state of -2

- **Hydrogen peroxide (H_2O_2)** 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 (MnO_2) to give water and oxygen:
  $2H_2O_2(l) \rightarrow 2H_2O + O_2(g)$
  In this reaction the O is both oxidised (from -1 to O, in O_2) and reduced (from -1 to -2, in H_2O)
  This reaction is a useful way of making oxygen gas in the laboratory

(v) Oxygen (O_2)

- Oxygen has a simple molecular structure; in its most common form it exists as a diatomic gas (O_2) and makes up 20% of the atmosphere in this form; it also exists as ozone (O_3), especially in the higher parts of the atmosphere (the stratosphere)

- Oxygen is prepared in the laboratory by the decomposition of hydrogen peroxide; it is prepared industrially by the fractional distillation of liquid air (the main components of air are N_2 (79%), O_2 (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:
  $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$
  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

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

- Fire is the glowing of gas molecules and dust particles as a result of them becoming very hot due to an exothermic reaction between a fuel and oxygen; this process is known as "burning":
  - a substance which can catch fire (burn) but only if energy is provided initially in the form of an ignition is said to be **flammable** (eg wood)
  - a substance which can catch fire spontaneously is said to be **inflammable** (eg white phosphorus)
  - a substance which cannot catch fire is said to be **non-flammable** (eg water or carbon dioxide)

- Fires are a very common way of disposing of unwanted material or creating heat; fires which are out of control, however, can be very dangerous; there are a number of different ways to put out a fire:
  - water removes heat from the fuel and therefore prevents it from burning; water can be used to put out fires caused by burning most flammable solids, such as wood; water should not be used for flammable liquids or gases, for electrical fires or for burning metals
  - carbon dioxide can put out fires by depriving the fuel of oxygen; it is best used for flammable liquids and gases, for electrical fires and for metals (water should not be used for these fires)
  - some types of non-flammable foam can also be used for flammable liquids and gases
  - all science laboratories and many households have fire extinguishers in order to put a fire out quickly; these usually contain carbon dioxide but they sometimes contain foam or water

Activity 6.2: What is the best way to put out a fire?
Watch this video on different fire extinguishers and their uses: www.youtube.com/watch?v=GjSoxJF3RD4

(vi) Oxides of other elements

- 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, O^{2-}; this is a base as it can accept protons from water to form OH^- ions: $O^{2-} + H_2O \rightarrow 2OH^-$; it can also react with acids to form water: $O^{2-} + 2H^+ \rightarrow H_2O$ (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): $CaO + H_2O \rightarrow Ca(OH)_2$; $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$

- 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: CO_2, SO_2 and SO_3
  - CO_2: dissolves in water to give a slightly acidic solution $CO_2 + H_2O \rightleftharpoons H^+ + HCO_3^-$; it reacts with hydroxides to form carbonates: $CO_2 + 2OH^- \rightarrow CO_3^{2-}$; when bubbled through limewater ($Ca(OH)_2$) it forms $CaCO_3$ and this is the basis of the test for CO_2 (see Unit 5 – Acids, Bases and Salts)
  - SO_2: dissolves in water to give an acidic solution $SO_2 + H_2O \rightleftharpoons 2H^+ + SO_3^{2-}$; SO_2 is sufficiently acidic to turn damp blue litmus paper red and this is the basis of the one of the tests for SO_2 (see Unit 5 – Acids, Bases and Salts)
  - SO_3: dissolves in water to give a strongly acidic solution: $SO_3 + H_2O \rightleftharpoons 2H^+ + SO_4^{2-}$; 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 H_2SO_4, HNO_3, $HClO$

- 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, Al_2O_3 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)

  $Al(OH)_3 + 3HCl \rightarrow AlCl_3 + 3H_2O$ and $Al(OH)_3 + NaOH \rightarrow NaAl(OH)_4$
  $ZnO + 2HCl \rightarrow ZnCl_2 + H_2O$ and $ZnO + 2NaOH \rightarrow Na_2ZnO_2 + H_2O$

- 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 6.3: Understanding the reactions of hydrogen, oxygen and their compounds

(a) Describe a simple test for hydrogen gas
(b) Draw the structure of hydrogen peroxide
(c) Write an equation for the decomposition of hydrogen peroxide and explain why this reaction is useful
(d) Outline the main stages in the industrial preparation of oxygen
(e) Write an equation for the reaction of glucose with oxgygen and hence explain why oxygen is essential for life
(f) What is the best way to put out a fire caused by burning wood?
(g) What is the best way to put out a fire caused by burning oil?
(h) Explain how you can predict whether an oxide or hydroxide will be acidic or basic
(i) Classify the following oxides and hydroxides as acidic, basic or amphoteric: MgO, NaOH, CO_2, SO_3, Al_2O_3, ClOH, $HPO(OH)_2$, FeO, SiO_2
(j) Write equations for the reactions of CaO with SO_2, NaOH with SiO_2 and CO_2 with MgO

Lesson 7 – What do I need to know about sulphur and its compounds?

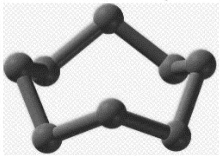
f) Sulphur and its Compounds

Summary Activity 7.1: 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
- Give the formula of sulphuric acid and write an equation to show its reaction with:
  (i) sodium hydroxide
  (ii) ammonia
  (iii) sodium chloride
  (iv) sodium bromide
  (v) 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

(i) Sulphur

- The element sulphur has a simple molecular structure; there are several allotropes of sulphur but most of these contain S_8 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 SO_2

- Sulphur also reacts with metals to make **sulphides**, which contain the S^{2-} ion:
  $Zn + S \rightarrow ZnS$
  $Cu + S \rightarrow CuS$

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

- **Sulphur dioxide (SO_2)** 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 **H_2SO_3 (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)
  - SO_2 also reacts with alkalis to make salts containing the **sulphate (IV) ion (SO_3^{2-})**:
    $SO_2 + 2OH^- \rightarrow SO_3^{2-}$

- **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); $SO_3^{2-} + 2H^+ \rightarrow SO_2 + H_2O$
  - this reaction can also be used in the laboratory reparation of sulphur dioxide

- 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**:
  $SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$;
  - 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 ^oC is used
  - a V_2O_5 catalyst is used in order to speed up the reaction and allow a lower temperature to be used

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

- **Sulphur trioxide (SO_3)** 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 (H_2SO_4)**: $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$
  - SO_3 is formed in the atmosphere when SO_2 in the air reacts with oxygen; acid rain is formed when SO_3 in the atmosphere dissolves in rainwater

- **Sulphuric acid (H_2SO_4)** is the acid produced when SO_3 is dissolved in water; it is produced on a huge scale industrially:

  - Sulphur is oxidised to sulphur dioxide ($S + O_2 \rightarrow SO_2$); Sulphur dioxide is oxidised to sulphur trioxide in the Contact Process: $SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$; 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 (SO_4^{2-})** (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):
    $2NH_3(g) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$

  - 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 HNO_3 and in the industrial preparation of phosphoric acid:
    $H_2SO_4 + 2KCl \rightarrow K_2SO_4 + 2HCl$
    $H_2SO_4 + 2KNO_3 \rightarrow K_2SO_4 + 2HNO_3$
    $3H_2SO_4 + Ca_3(PO_4)_2 + \rightarrow 3CaSO_4 + 2H_3PO_4$

  - Concentrated sulphuric acid is also an oxidising agent; it can oxidise bromide ions to bromine and iodide ions to iodine:
    $H_2SO_4 + 2Br^- + 2H^+ \rightarrow SO_2 + Br_2 + 2H_2O$ (S reduced from +6 to +4; Br oxidised from -1 to 0)
    $H_2SO_4 + 8I^- + 8H^+ \rightarrow H_2S + 4I_2 + 4H_2O$ (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:
    $Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$ (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 as an alternative laboratory preparation of sulphur dioxide gas

  - 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)

- **Sulphate (VI) ions (SO_4^{2-})** are formed when sulphuric acid reacts with bases; they are also formed when SO_2 or SO_3 ions are oxidised

  - the presence of SO_4^{2-} 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): $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

- 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

(iv) Sulphides (S^{2-})

- 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 S^{2-}

- All metal sulphides (like all metal oxides) are insoluble in water except for Group I sulphides such as Na_2S and K_2S and ammonium sulphide ($(NH_4)_2S$); the sulphide ion is strongly basic and reacts with water:

  $S^{2-} + H_2O \rightarrow HS^- + OH^-$; $HS^- + H_2O \rightleftharpoons H_2S + OH^-$

  Hydrogen sulphide gas (**H_2S**) is gradually evolved; H_2S smells of rotten eggs; H_2S is also formed when sulphide ions react with acids: eg $CuS(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2S(g)$

- Metal sulphides and hydrogen sulphide contain S in the -2 oxidation state; sulphides can be oxidised to SO_2 when they are heated in air: eg $2CuS + 3O_2 \rightarrow 2CuO + 2SO_2$

  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 7.2 – understanding sulphur and its compounds

(a) Explain why SO_2 and SO_3^{2-} ions are reducing agents and state a use for these species which relies on this property
(b) State and justify the conditions used in the conversion of SO_2 to SO_3 in the Contact process
(c) State the acid-base properties of sulphur trioxide and explain how acid rain is produced
(d) Give a use for sulphuric acid (i) as an acid; (ii) as an oxidising agent; (iii) as a dehydrating agent
(e) Write an equation to show how zinc reacts with sulphur
(f) Write an equation to show how zinc sulphide reacts with dilute hydrochloric acid
(g) Write an equation to show how zinc sulphide reacts with oxygen and explain why this reaction is useful
(h) State how the oxidation number of S changes during the reactions above
(i) State how hydrogen sulphide gas can be identified

Lesson 8 – What do I need to know about carbon and its inorganic compounds?

g) Carbon and its Compounds

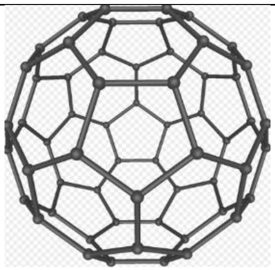
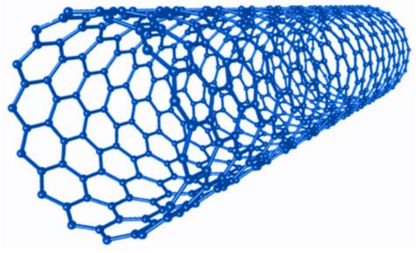
Summary Activity 8.1: What can you remember about carbon, carbon dioxide and carbonate ions?

- 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); most other allotropes of carbon exist in the shape of either balls or tubes and are known as **fullerenes;** carbon also exists in other, less pure forms, most notably **coal** and **coke**

- Fullerenes can be classified either as:
  - **Buckminster-Fullerenes**, which are large molecules of carbon atoms arranged in a spherical shape, similar to a football; the most common form is C_{60}
  - or as **nanotubes**, which are long cylinders of carbon atoms

Name	Bukminster-Fullerenes	Nanotubes
Diagram		
Properties	Cage-like structure which can trap other molecules	High surface area and good electrical conductivity
Uses	As molecular cages; they can trap free radicals produced during an allergic reaction and reduce inflammation; they can also trap hydrogen molecules and be used as a very small fuel tank	As catalysts or electrodes

- 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-butuminous 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

- 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 ^oC 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)

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

    $CO_2 + OH^- \rightarrow HCO_3^-$; $CO_2 + 2OH^- \rightarrow CO_3^{2-} + H_2O$

    Eg $CO_2 + NaOH \rightarrow NaHCO_3$; $CO_2 + 2NaOH \rightarrow Na_2CO_3 + H_2O$

    This reaction is the basis for the test for carbon dioxide: CO_2 reacts with limewater ($Ca(OH)_2$), 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):

    $CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$

    $CaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ca(HCO_3)_2(aq)$

  - Carbon dioxide is not flammable at all and is used in fire extinguishers; because it is denser 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)

    eg $MgO + CO_2 \rightarrow MgCO_3$
  - carbonate ions can also be formed in solution by dissolving carbon dioxide in alkalis

    $CO_2 + 2OH^- \rightarrow CO_3^{2-} + H_2O$
  - carbonate ions react with acids to form carbon dioxide: $CO_3^{2-} + 2H^+ \rightarrow CO_2 + H_2O$ (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 maxing standard solutions because it is stable in its most common form $Na_2CO_3.10H_2O$

  - **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; $CaCO_3 \rightarrow CaO + CO_2$; 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:

  $CO_2 + OH^- \rightarrow HCO_3^-$
  $CO_3^{2-} + H^+ \rightarrow HCO_3^-$

  - 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:

    $HCO_3^- + OH^- \rightarrow CO_3^{2-} + H_2O$
    $HCO_3^- + H^+ \rightarrow CO_2 + H_2O$;
    $2HCO_3^- \rightarrow CO_2 + CO_3^{2-} + H_2O$

  - **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: $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$; 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 tables to reduce stomach acidity (ie treat indigestion); it is also used to raise the pH of swimming pools

(iii) 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: 

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

    $2C + O_2 \rightarrow 2CO_2$
    $C + CO_2 \rightarrow 2CO$

  - 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 CO_2:

    $2CO + O \rightarrow 2CO_2$

Test your knowledge 8.2 – understanding carbon and its compounds

(a) Explain the difference between peat, lignite, coal and anthracite
(b) Explain how coke is produced from coal and list any other useful products of this process
(c) What is synthesis gas and how is it made?
(d) Write equations to show the acidic behaviour of carbon dioxide, the basic behaviour of carbonate ions and the amphoteric behaviour of hydrogencarbonate ions
(e) Describe the laboratory preparation of carbon dioxide and explain how the gas is collected
(f) Explain why carbon dioxide is used in fire extinguishers
(g) Give the formulae of sodium carbonate, calcium carbonate and sodium hydrogencarbonate and explain why each is useful

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

h) 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 (N_2)

- Nitrogen has a simple molecular structure (see Unit 2 – Particles, Bonding and Structure); it exists as a diatomic gas (N_2) 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)

- The main use of nitrogen is to manufacture **ammonia** (**NH_3**) by the **Haber process**:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g);$$

  - 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 ^{o}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 (NH_3) and ammonium ions (NH_4^+)

- **Ammonia (NH_3)** has a simple molecular structure; it exists as NH_3 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; $NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$; 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 it ability to neutralise HCl gas, forming a white smoke which is the salt NH_4Cl (see Topic 5 – Acids, Bases and Salts)

- Ammonia can also react with nitric acid and sulphuric acid to make the salts NH_4NO_3 and $(NH_4)_2SO_4$, which are both important fertilisers; this is the most important use of ammonia
  $NH_3 + HNO_3 \rightarrow NH_4NO_3$; $2NH_3 + H_2SO_4 \rightarrow (NH_2)_2SO_4$
  NH_4^+ ions are a very important source of nitrogen, which plants needed; bacteria in the soil gradually oxidise NH_4^+ ions to NO_3^- 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

- 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 $Ca(OH)_2$ is heated with solid NH_4Cl and the following reaction takes place: $Ca(OH)_2(s) + 2NH_4Cl(s) \rightarrow CaCl_2(s) + 2NH_3(g) + 2H_2O(l)$

- 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:
  $NaNO_2(s) + NH_4Cl(s) \rightarrow N_2(g) + NaCl(s) + 2H_2O(l)$
  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

- 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

Test your knowledge 9.1 – Understanding Nitrogen and Ammonia

(a) Outline the main stages in the industrial preparation of nitrogen
(b) Describe the laboratory preparation of nitrogen; state the type of reaction taking place and explain why the gas can be collected in this way
(c) 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
(d) Describe and justify the conditions used in the manufacture of ammonia by the Haber process
(e) State the main use of ammonia

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

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

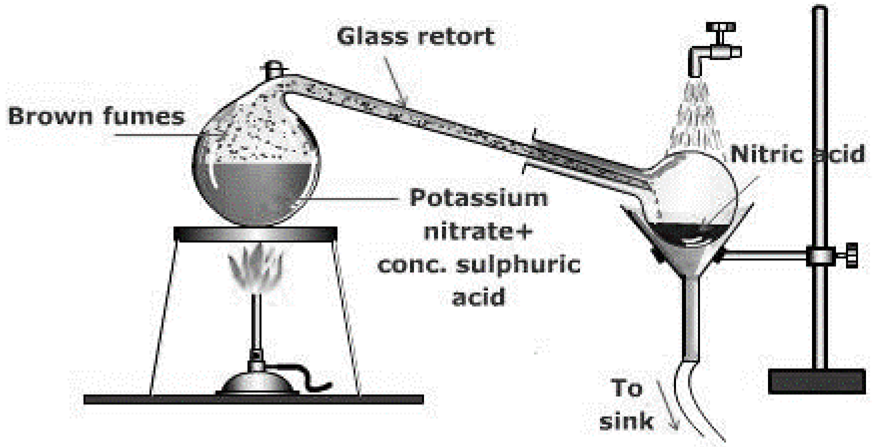
(iii) Nitric acid (HNO_3) and nitrate (V) ions (NO_3^-)

- **Hydrogen nitrate (HNO_3)** 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: $H^+(aq)$ + $NO_3^-(aq)$; 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 NO_3^- 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: $HNO_3 + NH_3 \rightarrow NH_4NO_3$

- 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:

  $H_2SO_4(l) + KNO_3(s) \rightarrow HNO_3(g) + K_2SO_4(s)$



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:
    $2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$
    The N is reduced from + 5 to +3 and the O is oxidised from -2 to 0
  - All other nitrates decompose to give oxides, $NO_2(g)$ and $O_2(g)$:
    $2Ca(NO_3)_2(s) \rightarrow 2CaO(s) + 4NO_2(g) + O_2(g)$
    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 – How can we prepare and collect gases in the laboratory? (Part 1)

i) The Preparation and Collection of Gases

Summary Activity 11.1: Which reactions can be used to prepare different gases in the laboratory?

- Go through the previous lessons on this topic and recall which chemical reactions are used in the laboratory preparation of the following gases: hydrogen (H_2), oxygen (O_2), nitrogen (N_2), chlorine (Cl_2), hydrogen chloride (HCl), sulphur dioxide (SO_2), carbon dioxide (CO_2) and ammonia (NH_3); in each case, write an equation for the reaction and state the type of reaction occurring
- State the chemical tests for all the above gases

(i) General Principles of Collection of Gases

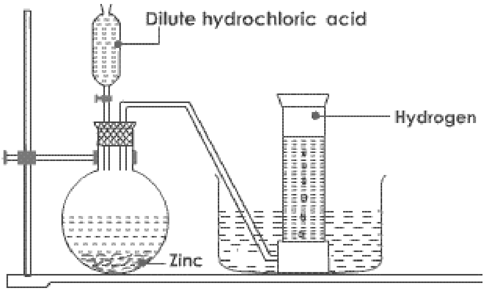
- Earlier in this topic you have studied the chemical reactions used to prepare the following gases in the laboratory: hydrogen (H_2), oxygen (O_2), nitrogen (N_2), chlorine (Cl_2), hydrogen chloride (HCl), sulphur dioxide (SO_2), carbon dioxide (CO_2) and ammonia (NH_3)

- When the gas is produced it needs to be collected; gases can be collected in one of three ways:
  - by upward delivery over water; many gases can be conveniently collected by bubbling the gas through water and into an inverted gas jar, cylinder or boiling tube which was full of water; the gas displaces the water and the container fills up with the gas
  - by upward delivery without water; in some cases, the gas is released **upwards** into an inverted container which is not full of water (ie which just contains air)
  - by downward delivery without water; in some cases, the gas is released **downwards** into a container which is not full of water (ie which just contains air)

- The best method for collecting a particular gas depends on two properties of the gas:
  - its solubility in water: gases which are insoluble or only slightly soluble in water are always collected by upward delivery over water, as this is the most efficient way to collect a gas; gases which are highly soluble in water, however, cannot be collected by this method (eg HCl and NH_3) and have to be collected without water
  - its density: the density of a gas is directly proportional to its molecular mass; the average molecular mass of air is 28.4; gases with a molecular mass of less than 28 are less dense than air and move upwards through air; these gases are collected by upward delivery ; gases with a molecular mass of more than 28 are denser than air and move downwards through air; these gases are collected by downward delivery; note that if a gas is insoluble in water it can always be collected over water, regardless of its density

(ii) hydrogen

- **Hydrogen (H_2)** is prepared in the laboratory by reacting dilute hydrochloric acid with a fairly reactive metal such as zinc
  - it is a redox reaction (metal-acid): $Zn + 2HCl \rightarrow Zn + ZnCl_2 + H_2$ (Zn oxidised from 0 to +2, H reduced from +1 to 0)
  - hydrogen is insoluble in water as it is non-polar; it can therefore be collected over water:



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Demonstration 11.2: 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)

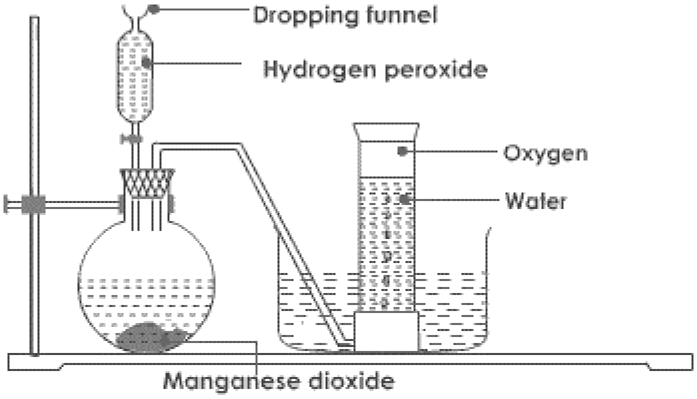
1) Fill the dropping funnel with 2 $moldm^{-3}$ HCl; add around 20 zinc granules into the flask
2) Gradually add the HCl to the zinc; you will observe a gas being given off which will collect in the gas jar
3) Once the gas jar is almost full, slide a lid under the gas jar, remove it from the water; keep the lid of the gas jar under the jar
4) 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

(iii) oxygen

- **Oxygen (O_2)** is prepared in the laboratory by the decomposition of hydrogen peroxide in the presence of a catalyst (MnO_2)
  - it is a redox reaction (disproportionation): $2H_2O_2 \rightarrow 2H_2O + O_2$ (O is oxidised from -1 and reduced from -1 to -2)
  - oxygen has limited solubility in water as it is non-polar, so it can be collected over water:



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Demonstration 11.3: Prepare O_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)

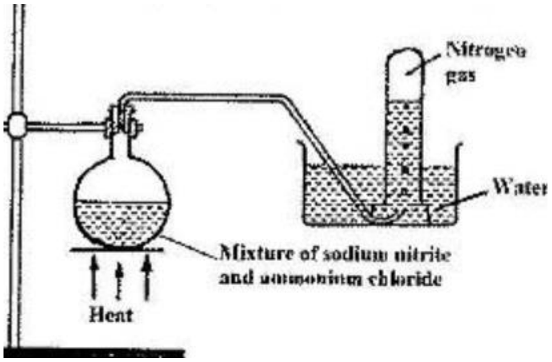
1) Add around 100 cm^3 of water and one spatula of MnO_2 into the flask; fill the dropping funnel with concentrated H_2O_2 (2 $moldm^{-3}$)
2) Gradually add the H_2O_2 to the flask; you will observe a gas being given off which will collect in the gas jar
3) Once the gas jar is almost full, slide a lid under the gas jar and remove it from the water
4) Confirm that the gas is O_2 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 O_2 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

(iv) Nitrogen

- **Nitrogen (N_2)** is prepared in the laboratory by the reaction of ammonium chloride with sodium nitrate (III)
  - it is a redox reaction: $NH_4Cl + NaNO_2 \rightarrow N_2 + 2H_2O + NaCl$ (N in NH_4^+ is oxidised from -3 to 0 and N in NO_2^- is reduced from +3 to 0)
  - nitrogen has limited solubility in water as it is non-polar, so it can be collected over water:



Demonstration 11.4: Prepare N_2 gas in the laboratory

Set up the apparatus as shown in the diagram above (a flask, a delivery tube, a bung to fit the flask with a hole for a delivery tube, a Bunsen burner with tripod and gauze, a trough of water and a gas jar)

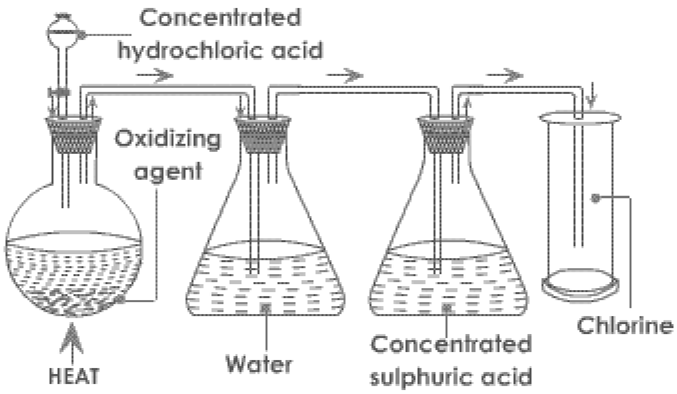
1) Dissolve 14 g $NaNO_2$ and 11 g of NH_4Cl in around 200 cm^3 of water in the round-bottomed flask
2) Heat the mixture gently; you will observe a gas being given off which will collect in the gas jar
3) Once the gas jar is almost full, slide a lid under the gas jar and remove it from the water

Write an equation for the reaction taking place
What is oxidised and what is reduced?
Why can N_2 be collected over water?

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=6mos8zB36TY

(v) chlorine

- **Chlorine (Cl_2)** is prepared in the laboratory by the reaction of MnO_2 with concentrated HCl
  - this is a redox reaction: $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$ (Mn reduced from +4 to +2; Cl^- oxidised from -1 to 0)
  - the chlorine should be bubbled through water first (to remove HCl) and then through concentrated sulphuric acid (to remove water)
  - chlorine reacts with water so should not be collected over water
  - a solution of $KMnO_4$ can be used as an alternative to MnO_2
  - chlorine is denser than air ($m_r = 71$) so should be collected by downward delivery:



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Demonstration 11.5: 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 bungs to fit the flasks, each with two holes, thee delivery tubes, a gas jar, Bunsen burner, tripod, gauze)
2) Fill the dropping funnel with concentrated HCl; 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 and start to heat the mixture gently; 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 (note that this preparation uses a different oxidising agent)

Test your knowledge 11.6: Describing the collection of different gases (Part 1)

(a) Write an equation for the reactions used to prepare the following gases in the laboratory:
  (i) hydrogen
  (ii) oxygen
  (iii) nitrogen
  (iv) chlorine

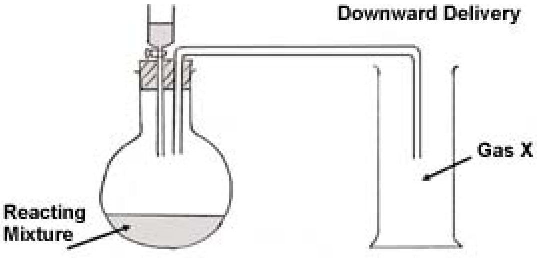
(b) For each of the gases above, state how the gas is collected and explain any differences in the way the gases are collected

(c) Describe a simple chemical test which could confirm the identify of:
  (i) hydrogen
  (ii) oxygen
  (iii) chlorine

Lesson 12 – How can we prepare and collect gases in the laboratory? (Part 2)

(vi) Hydrogen chloride

- **Hydrogen chloride (HCl)** is prepared in the laboratory by reaction of concentrated H_2SO_4 with solid KCl
  - this is an acid-salt reaction: $H_2SO_4 + 2KCl \rightarrow K_2SO_4 + 2HCl$
  - HCl is very soluble in water so cannot be collected over water
  - because HCl is denser than air ($m_r = 36.5$) it is collected by downward delivery:



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

Demonstration 12.1: Prepare HCl gas in the laboratory

(CAUTION – HCl gas is highly toxic – use a fume cupboard; concentrated H_2SO_4 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, clamp, stand and boss)
2) Add around 5 cm^3 of concentrated H_2SO_4 into the dropping funnel; add 2 spatulas 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 gar 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 NH_3)

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 NH_3

(vii) Sulphur dioxide

- **Sulphur dioxide (SO_2)** can be prepared in the laboratory by in one of two ways:
  - (i) by the addition of dilute HCl to a solid sample of Na_2SO_3; this is an acid-base reaction:
    $Na_2SO_3 + 2HCl \rightarrow 2NaCl + SO_2 + H_2O$
  - (ii) by the reduction of concentrated sulphuric acid using copper metal; this is a redox reaction:
    $Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$ (Cu oxidised from 0 to +2; S reduced from +6 to +4)
  - sulphur dioxide cannot be collected over water as it is soluble in water (it is polar)
  - the gas can be passed through concentrated H_2SO_4 in order to remove any water vapour
  - sulphur dioxide is denser than air (m_r = 64) so it should be collected by downward delivery:

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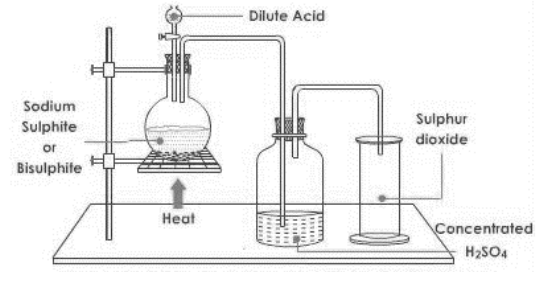
<td>

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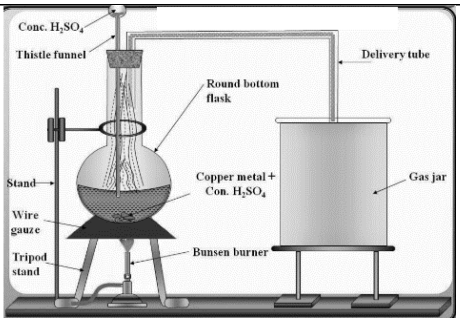
method (i)</td>
<td>

www.youtube.com
method (ii)</td>
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Demonstration 12.2: Prepare SO_2 gas in the laboratory (Part 1)

1) Set up the apparatus as shown in the method (i) diagram above (a round-bottomed flask, a dropping funnel, two delivery tubes, conical flask two bungs to fit the flasks, each with two holes, a gas jar, Bunsen, tripod, gauze, clamp, boss, stand)
2) Fill the dropping funnel with 1.0 $moldm^{-3}$ HCl; add around 5 spatulas 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 stare 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

Demonstration 12.3: Prepare SO_2 gas in the laboratory (Part 2)

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

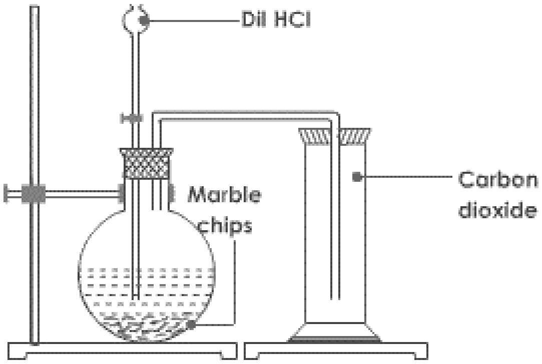
1) Set up the apparatus as shown in the diagram above (a round-bottomed flask, a dropping funnel, a delivery tube, a bung to fit the flask with two holes, a gas jar, a Bunsen burner with tripod and gauze, a clamp with stand and boss)
2) Fill the dropping funnel with concentrated H_2SO_4; add around 5 spatulas 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 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 is oxidised and what is reduced?

Can't do this experiment? Watch this video instead: www.youtube.com/watch?v=au04TpZOON0 (go to 4:15)

(viii) Carbon dioxide

- **carbon dioxide (CO_2)** is generally prepared in the laboratory by reacting carbonates with acids
  - this is an acid-base reaction: $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2 + H_2O$
  - carbon dioxide is fairly soluble in water so should not be collected over water (although it is possible to collect some gas by the method)
  - carbon dioxide is denser than air ($m_r = 44$) so is best collected by downward delivery



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Demonstration 12.4: Prepare CO_2 gas in the laboratory

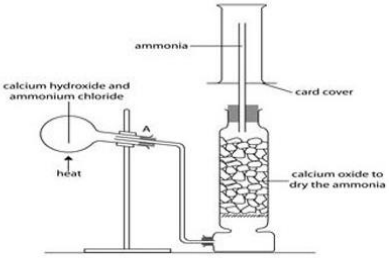
1) Set up the apparatus as shown in the diagram above (a round-bottomed flask, a dropping funnel, a delivery tube, a bung to fit the flask with two holes, a gas jar, a clamp with stand and boss)
2) Fill the dropping funnel with 2 $moldm^{-3}$ HCl; add around 5 spatulas of marble chips into the round-bottomed flask
3) Whilst the gas is being produced, dip the delivery tube into a test tube containing limewater 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 CO_2 have on limewater and why?
What effect does CO_2 have on a lit splint and why?

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

(ix) Ammonia

- **ammonia (NH_3)** is generally prepared in the laboratory by warming a solid hydroxide with a solid sample of an ammonium salt
  - this is an acid-base reaction: $Ca(OH)_2(s) + 2NH_4Cl(s) \rightarrow CaCl_2(s) + 2NH_3(g) + 2H_2O(l)$
  - ammonia is very soluble in water so cannot be collected over water
  - ammonia is less dense than air ($m_r = 17$) so is best collected by upward delivery
  - the NH_3 produced also contains water vapour and should be passed through a drying agent such as CaO before it is collected



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Demonstration 12.5: Prepare NH_3 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 CaO tower and a gas jar, and a Bunsen burner)

1) Add two spatulas each of ammonium chloride and calcium hydroxide to the boiling tube; shake the mixture so it is well mixed
2) 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
3) Test the gas by placing damp red 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

Test your knowledge 12.6: Describing the collection of different gases (Part 2)

(a) Write an equation for the reactions used to prepare the following gases in the laboratory:
  (i) hydrogen chloride
  (ii) sulphur dioxide
  (iii) carbon dioxide
  (iv) ammonia
(b) For each of the gases above, state how the gas is collected and explain any differences in the way the gases are collected
(c) Describe a simple chemical test which could confirm the identify of:
  (i) hydrogen chloride
  (ii) sulphur dioxide
  (iii) carbon dioxide
  (iv) ammonia

Lesson 13 – What do I need to know about rocks?

j) 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
    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 13.1 – understanding rocks

(a) Explain the difference between a rock and a mineral
(b) Describe the three types of rock and explain how each is formed
(c) Explain what is meant by the term "weathering"
(d) Give three examples of chemical weathering

Lesson 14 – What have I learned about non-metals and their compounds?

14.1 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, NH_3 and CO_2; 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.