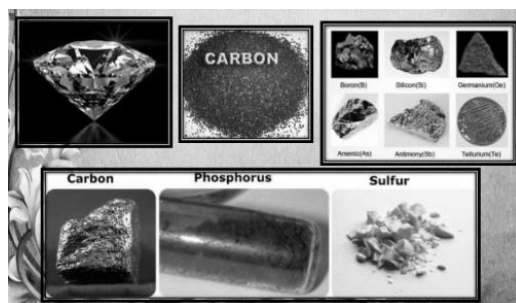


UNIT 12

NON-METALS AND THEIR COMPOUNDS

Teacher Version



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Contents

- (a) Introduction to Non-Metals
- (b) The Noble Gases
- (c) The Halogens and their Compounds
- (d) Water
- (e) Hydrogen, Oxygen and their Other Compounds
- (f) Sulphur and its Compounds
- (g) Carbon and its Compounds
- (h) Nitrogen and its Compounds
- (i) Laboratory Preparation and Collection of Gases
- (j) Rocks

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

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

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

Estimated Teaching Time: 14 hours

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

UNIT 12 SUMMARY AND SYLLABUS REFERENCE

Lesson	Title and Syllabus Reference
1	<p>Introduction to Non-Metals; Noble Gases; Physical Properties of Halogens CA2bi periodic gradation of elements in group i.e. the halogens: F₂/Cl₂/Br₂/I₂ (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)</p>
2	<p>Chemical Properties of Halogens and Halides I CA2bi periodic gradation of elements in group i.e. the halogens: F₂/Cl₂/Br₂/I₂ (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)</p>
3	<p>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)</p>
4	<p>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)</p>
5	<p>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)</p>
6	<p>Hydrogen, Oxygen and their other Compounds CA5bii the gases - properties of gases (chemical properties of H₂); 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 SO₃²⁻ and CO₃²⁻ (no formal scheme of analysis is required); CA13civ characteristic test for H₂ (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 H₂ gas); ISA10.2 uses of non-metals (uses of O₂); ISE6.2 first aid methods (demonstration of the following using models: methods of extinguishing different fires)</p>
7	<p>Sulphur and its Compounds CA5bii the gases – preparation and properties of gases (chemical properties of SO₂); 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 SO₃²⁻, confirmatory test for SO₃²⁻ and of conc. H₂SO₄ on solid samples of Cl⁻ (no formal scheme of analysis is required); CA13civ characteristic test for SO₂ gas (no formal scheme of analysis is required); CC7ei Sulphur – allotropes and uses; CC7eii sulphur – compounds of sulphur: sulphides; CC7eiii 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)</p>

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

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 CO₃²⁻, confirmatory test for CO₃²⁻ (no formal scheme of analysis is required); CA13civ characteristic test for CO₂ 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 ₂ CO 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 NH₄⁺ with NaOH, confirmatory test for NH₄⁺ (no formal scheme of analysis is required); CA13civ characteristic test for NH₃ 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. H₂SO₄ on solid samples of NO₃⁻; confirmatory test for NO₃⁻ (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

1	2	3	4	5	6	7	0										
(1) 6.9 Li lithium 3	(2) 9.0 Be beryllium 4	(3) 23.0 Na sodium 11	(4) 47.9 Ti titanium 22	(5) 50.9 V vanadium 23	(6) 52.0 Cr chromium 24	(7) 54.9 Mn manganese 25	(8) 55.8 Fe iron 26	(9) 58.9 Co cobalt 27	(10) 58.7 Ni nickel 28	(11) 63.5 Cu copper 29	(12) 65.4 Zn zinc 30	(13) 10.8 B boron 5	(14) 12.0 C carbon 6	(15) 14.0 N nitrogen 7	(16) 16.0 O oxygen 8	(17) 19.0 F fluorine 9	(18) 4.0 He helium 2
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	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
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] 98.0 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	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	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			

1.0
H
hydrogen
1

Key
relative atomic mass
symbol
name
atomic (proton) number

*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
- A non-metal is an element which does not contain metallic bonding (or which contains covalent bonding)
 - Hydrogen, carbon, oxygen, nitrogen, fluorine, chlorine, bromine, iodine
 - Diamond - giant covalent structure (lattice of carbon atoms held together by covalent bonds)
 - Chlorine – two atoms held together by a covalent bond to form a molecule; weak Van der Waal's forces between the molecules

- **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 H₂, O₂, N₂), 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

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- The noble gases all have very low boiling points due to the very weak Van der Waal's forces between the atoms; on descending the Group from He to Xe the number of electrons per atom increases and so the Van der Waal's forces become stronger and the boiling point increases; He boils at 4 K (-269 °C) and Xe boils at 166 K (-107 °C)
- The density of the noble gases increases down the group as the mass of the atoms increase; He and Ne are less dense than air; Ar, Kr and Xe are 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: Understanding Group 0 – the noble gases

- (a) Explain why the noble gases are so unreactive
 - (b) State and explain the trend in boiling points of the noble gases
 - (c) State and explain the trend in densities of the noble gases
 - (d) State a use for each of helium, neon and argon
- (a) Eg ionisation energies very high so cannot lose electrons, outer shell full so cannot share or accept electrons easily
 - (b) Increase on descending the group; more electrons per atom so stronger Van der Waal's forces between atoms
 - (c) Increase on descending the group; mass of nucleus increases much faster than volume of atoms
 - (d) Helium – hot air balloons/oxygen tanks for diving; neon – lighting; argon – inert atmosphere

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 F₂, Cl₂, Br₂ and I₂; 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

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- 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: Understanding Group 7 – the halogens (physical properties)

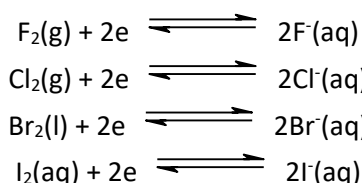
- State the colour of pure iodine, iodine in cyclohexane and iodine in water
- State the colours of chlorine and bromine when dissolved in cyclohexane or water
- State and explain the trend in boiling points of the halogens
- The final halogen is astatine; suggest whether it will be a solid, liquid or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

- Eg ionisation energies very high so cannot lose electrons, outer shell full so cannot share or accept electrons easily
- Increase on descending the group; more electrons per atom so stronger Van der Waal's forces between atoms
- Increase on descending the group; mass of nucleus increases much faster than volume of atoms
- Helium – hot air balloons/oxygen tanks for diving; neon – lighting; argon – inert atmosphere

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):

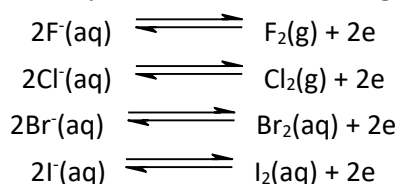


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- They are most commonly found in the -1 oxidation state, as X^- ions; these are known as **halide** ions
- Halide ions have no colour – they appear as colourless in aqueous solution and as white in the solid state
- Fluorine is the strongest oxidising 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

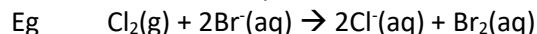
- The halide ions are reducing agents, as they can lose electrons and get oxidised (see Unit 6 – Redox Reactions):



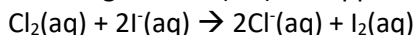
Iodide ions are the most reducing, followed by bromide ions; fluoride ions have no significant reducing properties; as the number of shells in the ion increases, there is more shielding of the nucleus and the outer electrons become less strongly held; these electrons are thus lost more easily and the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

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

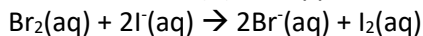
- The displacement reactions of halogens with halide ions provides a clear illustration of the trends in oxidizing properties of the halogens and the trends in reducing properties of the halide ions in aqueous solution; the more reactive halogens (ie the strongest oxidising agents) will displace the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



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



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



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

UNIT 12 – NON-METALS AND THEIR COMPOUNDS



Test your knowledge 2.1: Understanding Group 7 – 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
-
- (a) Decreases down the group from fluorine to iodine; more shells means more shielding, so electrons are less strongly attracted into the outer shell on descending the group and the halogen does not accept an electron as readily
 - (b) Increases down the group from fluoride to iodide; more shells means more shielding, so electrons in the outer shell are less strongly held and can be more easily lost
 - (c) The orange colour disappears and a yellow/brown colour appears; this is because I_2 is being produced; $Br_2 + 2I^- \rightarrow 2Br^- + I_2$; Br_2 is a stronger oxidising agent than I_2 so Br will displace I from its compounds
 - (d) No reaction; Br_2 is a weaker oxidising agent than I_2 so Br will not displace Cl from its compounds

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Practical 2.2: Carry out halogen displacement reactions

Note: take care not to inhale the fumes of Cl₂ or Br₂; 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 cm³ of KCl(aq) into two separate test tubes
- 2) Add 2 cm³ of bromine water to one of the test tubes and 2 cm³ of iodine water to the other test tube and record your observations
- 3) Add 1 cm³ of cyclohexane to both test tubes and shake well; record your observations
- 4) Pour 2 cm³ of KBr(aq) into two separate test tubes
- 5) Add 2 cm³ of chlorine water to one of the test tubes and 2 cm³ of iodine water to the other test tube and record your observations
- 6) Add 1 cm³ of cyclohexane to both test tubes and shake well; record your observations
- 7) Pour 2 cm³ of KI(aq) into two separate test tubes
- 8) Add 2 cm³ of chlorine water to one of the test tubes and 2 cm³ of bromine water to the other test tube and record your observations
- 9) Add 1 cm³ of cyclohexane to both test tubes and shake well; record your observations
- 10) Summarise your observations in a table similar to the one below:

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

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

Why do some of the reactions not take place?

Chemicals needed per group: access to bottles containing solutions of Cl₂, Br₂, I₂ (max conc. Possible), KCl, KBr, KI (1 M) and cyclohexane - each group will require around 5 cm³ - each bottle must have its own dropping pipette

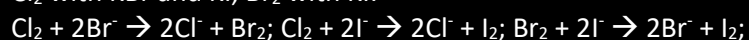
Apparatus needed per group: 6 test tubes and a test tube rack

In water, all solutions can appear yellow/orange/brown after the reaction; the addition of cyclohexane should give a distinct orange colour (for bromine) and a distinct purple colour (for iodine)

	Cl ₂		Br ₂		I ₂	
	water	cyclohexane	water	cyclohexane	water	cyclohexane
Cl ⁻			yellow/orange	orange	yellow/brown	purple
Br ⁻	yellow/orange	orange			yellow/brown	purple
I ⁻	yellow/brown	purple	yellow/brown	purple		

The halogen changes during the following reactions:

Cl₂ with KBr and KI, Br₂ with KI:



UNIT 12 – NON-METALS AND THEIR COMPOUNDS

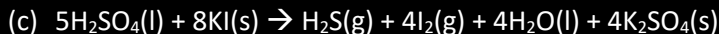
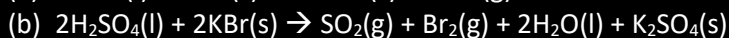
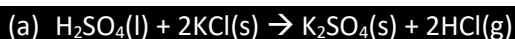
Lesson 3 – How else can we compare the reducing properties of halide ions?

- 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
$$2\text{H}_2\text{SO}_4(\text{l}) + 2\text{KBr}(\text{s}) \rightarrow \text{SO}_2(\text{g}) + \text{Br}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + \text{K}_2\text{SO}_4(\text{s})$$
this is a redox reaction; Br^- is oxidised to Br_2 ; Br_2 can be seen as an orange vapour; SO_2 can be tested for using dichromate paper (see Unit 6 – Redox Reactions)
$$5\text{H}_2\text{SO}_4(\text{l}) + 8\text{KI}(\text{s}) \rightarrow \text{H}_2\text{S}(\text{g}) + 4\text{I}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l}) + 4\text{K}_2\text{SO}_4(\text{s})$$
this is a redox reaction; I^- is oxidised to I_2 ; I_2 can be seen as a purple vapour
 - Cl^- is a much weaker reducing agent so is not oxidised by concentrated sulphuric acid; an acid-base reaction takes place and HCl is produced:
$$\text{H}_2\text{SO}_4(\text{l}) + 2\text{KCl}(\text{s}) \rightarrow \text{K}_2\text{SO}_4(\text{s}) + 2\text{HCl}(\text{g})$$
this is not a redox reaction; white fumes of HCl are seen; **this is the standard laboratory method for the preparation of HCl gas**
- It is possible to oxidise chloride ions to chlorine in the presence of a very strong oxidising agent, such as MnO_2 :
$$\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + 2\text{H}_2\text{O}$$
 - the Mn is reduced from +4 to +2; the Cl^- is oxidised to Cl_2 (-1 to 0)
 - **this is the standard laboratory preparation of Cl_2 gas**



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

- Write an equation for the reaction of potassium chloride with concentrated sulphuric acid
- Write an equation for the reaction of potassium bromide with concentrated sulphuric acid
- Write an equation for the reaction of potassium iodide with concentrated sulphuric acid
- Explain how the equations in (g), (h) and (i) illustrate the trend in reducing properties of the halides
- 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?



(d) Cl^- is the weakest reducing agent; it does not reduce H_2SO_4 ; Br^- is a stronger reducing agent; it reduces S in H_2SO_4 from +6 to +4; I^- is the strongest reducing agent; it reduces S in H_2SO_4 from +6 to -2

(e) MnO_2 ; it should also oxidise Br^- and I^- because these are stronger reducing agents than Cl^-

UNIT 12 – NON-METALS AND THEIR COMPOUNDS



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 H₂SO₄ is highly corrosive; wear gloves)

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

Equipment and chemicals needed: three test tubes, 2 g NaCl, 2 g KBr, 2 g KI, each with a spatula, one test tube rack; 10 cm³ of conc H₂SO₄ with dropping pipette; filter paper soaked in conc NH₃, dichromate paper - you must use gloves and a fume cupboard for this experiment

The vapour with KCl is white; the vapour with KBr is orange; the vapour with KI is purple; NH₃ will give a white smoke with all gases as HBr and HI are also produced and also give a white smoke with ammonia

White – HCl; orange – Br₂; purple – I₂; KCl reaction not redox; in KBr reaction S in H₂SO₄ is reduced and Br⁻ is oxidised; in KI reaction S in H₂SO₄ is reduced and I⁻ is oxidised

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?

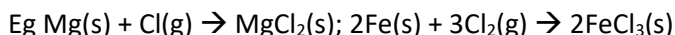


Summary Activity 4.1: What can you remember about tests for halide ions?

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

- Add HNO₃ and then AgNO₃; Cl⁻ will give a white precipitate, Br⁻ will give a cream precipitate and I⁻ will give a yellow precipitate

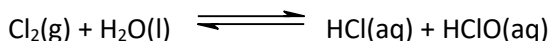
- Chlorine will combine directly with most metals to make metal chlorides; the chlorine is reduced (from 0 to -1) and the metal is oxidised:



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

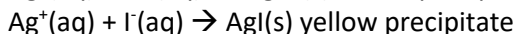
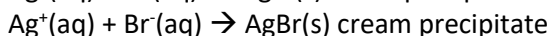
UNIT 12 – NON-METALS AND THEIR COMPOUNDS

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



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

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



(iii) Uses of Halogens and Halides

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



Test your knowledge 4.2: Group 7 – describing other reactions of halogens and halides

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

- $\text{Mg} + \text{Cl}_2 \rightarrow \text{MgCl}_2$; $\text{Mg} + \text{Br}_2 \rightarrow \text{MgBr}_2$
- Chlorine is more reactive than bromine as it attracts electrons more strongly
- $\text{Cl}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCl}(\text{aq}) + \text{HClO}(\text{aq})$; HClO is an oxidising agent and can kill bacteria but is not harmful to humans in small concentrations
- $\text{Cl}_2(\text{g}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{NaClO}(\text{aq}) + \text{H}_2\text{O}(\text{l})$; NaClO is a strong oxidising agent and is used in domestic bleach
- It changes from 0 to +1 and -1
Use blue litmus paper; chlorine turns it red and then bleaches it
- Add HNO_3 and then AgNO_3 ; Cl^- will give a white precipitate, Br^- will give a cream precipitate and I^- will give a yellow precipitate
- Chlorine used to sterilise water; iodine used to sterilise wounds; sodium chlorate (I) used in bleach, silver bromide used in photography

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?
- simple molecular; water forms molecules (H_2O) which are held together by hydrogen bonding
 - Heat a sample of water in a flask connected to a distillation apparatus with a thermometer in the head; record the maximum temperature reached; the intermolecular forces in water are stronger than between other molecules of a similar size, due to hydrogen bonding
 - Most ionic compounds dissolve in water, especially those containing NH_4^+ , K^+ , Na^+ or NO_3^- ; simple molecules dissolve in water only if they are polar or can form hydrogen bonds; most giant covalent substances do not dissolve in water

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

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

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

Apparatus needed: an inflated balloon and some water

The water stream should bend towards the balloon

This is because water is polar and polar molecules are attracted to charged particles

Ethanol and ethanoic acid should give the same effect

Most oils and paraffin should give no 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 cm³ of water into each test tube
- 2) Into the first test tube, place a spatula of salt; into the second, place a spatula of sugar; into the third, a piece of chalk; into the fourth, a spatula of sand; into the fifth, a piece of wax; into the sixth, 2 cm³ of ethanol and into the seventh, 2 cm³ of paraffin
- 3) Shake each test tube well for a few minutes and then allow them to stand for a few minutes; note your observations

Which substances are soluble in water?

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

Chemicals needed: salt, sugar, chalk, wax, sand (1 g per group), each container needs its own spatula; paraffin and ethanol (2 cm³ per group) each bottle needs its own dropping pipette, bottles of distilled water

Apparatus needed per group: seven test tubes and one test tube rack; one 10 cm³ measuring cylinder

Salt, sugar and ethanol will dissolve; wax, sand, chalk and paraffin will not

- Salt and chalk are ionic; ionic compounds sometimes dissolve in water, but only if the attraction between the ions and water is stronger than the attraction between the ions; in NaCl it is but in CaCO₃ it is not
- wax, sugar, ethanol and paraffin are simple molecular; simple molecular substances will dissolve in water if they can form hydrogen bonds with water; sugar and ethanol (which have -OH groups) can, wax and paraffin (which are non polar) cannot
- sand is giant covalent and giant covalent structures are usually insoluble in water

- 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 Cu²⁺, which forms a blue complex ion with water; anhydrous copper sulphate turns blue when water is added and **this is a test for water** (see Unit 9 – Metals and their Compounds)
- Water is the most useful substance on earth; it is the basis for human life as the human body relies on water to transport minerals and nutrients around the body and as a solvent for chemical reactions in cells; water is also used for washing; water is used as an industrial coolant, especially in power stations; water is also used as a raw material for making other useful substances, such as ethanol

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

- 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
- (a) High boiling point, high surface tension and lower density in solid state than liquid state; these anomalous properties are due to strong hydrogen bonding in water
 - (b) Sodium chloride (and most ionic compounds), sucrose, ethanol (and most molecules with hydrogen bonding) all dissolve in water; chalk (and some other ionic compounds), sand (and most giant covalent substances), wax and paraffin (and most non-polar molecules) do not
 - (c) Add NaOH or AgNO₃ - a precipitate may form if dissolved substances are present; test pH - if not 7 then dissolved substances are present; test conductivity - if conducts then dissolved substances are present; evaporate off water – if residue left then dissolved substances are present
 - (d) Add a few drops to anhydrous copper sulphate; if the copper sulphate turns blue then water is present
 - (e) Add positively charged ions to water, which causes some particles and dissolved substances to coagulate; separate off larger particles which settle at bottom; filter remaining water; add disinfectant

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

- both simple molecular; both form diatomic molecules with very weak Van der Waal's forces holding the molecules together
- Eg $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$; Mg, Al, Zn, Fe, Sn and Pb react steadily with acids; hydrogen gas is produced in this reaction
- H^+ ions turn blue litmus red; also it will give bubbles if added to CaCO_3 or another carbonate
- OH^- ions turn red litmus paper blue; they also give a pungent-smelling gas when warmed with ammonium chloride
- CaO : $\text{CaO} + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2\text{O}$; Na_2O : $\text{Na}_2\text{O} + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{H}_2\text{O}$
- SO_2 : $\text{SO}_2 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_3 + \text{H}_2\text{O}$; Cl_2O : $\text{Cl}_2\text{O} + 2\text{NaOH} \rightarrow 2\text{NaClO} + \text{H}_2\text{O}$
- ZnO : $\text{ZnO} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2\text{O}$; $\text{ZnO} + 2\text{NaOH} \rightarrow \text{Na}_2\text{ZnO}_2 + \text{H}_2\text{O}$

(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 $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$
- 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):
 $\text{Zn(s)} + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g})$ or $\text{Zn(s)} + 2\text{HCl(aq)} \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g})$
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 $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$ (see Unit 5 – Acids, Bases and Salts)
- Hydroxide ions turn red litmus blue; if warmed with ammonium chloride, hydroxide ions will give off ammonia gas which has a characteristic pungent smell (see Unit 5 – Acids, Bases and Salts)

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

(iv) Hydrogen Peroxide (H₂O₂)

- The most common oxide of hydrogen is water (H₂O); this contains oxygen in an oxidation state of -2
- **Hydrogen peroxide (H₂O₂)** is a less common oxide of hydrogen; it contains O in an oxidation state of -1; all compounds containing O in an oxidation state of -1 are called peroxides
- Hydrogen peroxide is a useful oxidising agent known for its bleaching properties; it is especially used in the bleaching of hair
- Hydrogen peroxide decomposes in the presence of a catalyst (MnO₂) to give water and oxygen:
$$2\text{H}_2\text{O}_2(\text{l}) \rightarrow 2\text{H}_2\text{O} + \text{O}_2(\text{g})$$

In this reaction the O is both oxidised (from -1 to 0, in O₂) and reduced (from -1 to -2, in H₂O)

This reaction is a useful way of making oxygen gas in the laboratory

(v) Oxygen (O₂)

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

This exothermic reaction is the source of energy all plants and animals need to live; oxygen can also react with other fuels, mostly hydrocarbons or alcohols; these are exothermic reactions which produce energy
- 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)

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

- 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?

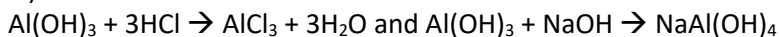
Watch this video on different fire extinguishers and their uses: www.youtube.com/watch?v=GjSoxJF3RD4

(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 , $HClO$

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

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



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



Test your knowledge 6.3: Understanding the reactions of hydrogen, oxygen and their compounds

- Describe a simple test for hydrogen gas
- Draw the structure of hydrogen peroxide
- Write an equation for the decomposition of hydrogen peroxide and explain why this reaction is useful
- Outline the main stages in the industrial preparation of oxygen
- Write an equation for the reaction of glucose with oxygen and hence explain why oxygen is essential for life
- What is the best way to put out a fire caused by burning wood?
- What is the best way to put out a fire caused by burning oil?
- Explain how you can predict whether an oxide or hydroxide will be acidic or basic
- Classify the following oxides and hydroxides as acidic, basic or amphoteric: MgO, NaOH, CO₂, SO₃, Al₂O₃, ClOH, HPO(OH)₂, FeO, SiO₂
- Write equations for the reactions of CaO with SO₂, NaOH with SiO₂ and CO₂ with MgO

- Insert a lit splint into the test tube – the hydrogen burns with a squeaky pop
- H-O-O-H
- $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$; used in the laboratory preparation of oxygen
- The air is cooled to a very low temperature to ensure that N₂, O₂ and Ar all condense; liquid air run into base of fractionating column; temperature allowed to rise slowly; N₂ and Ar boil first, leaving oxygen at base of column
- $\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O}$; this is how all living things generate energy to function
- Water
- A fire extinguisher containing carbon dioxide or foam
- Covalent oxides (ie oxides of non-metals) are acidic; ionic oxides (ie oxides of metals) are basic
- Acidic: CO₂, SO₃, ClOH, HPO(OH)₂, SiO₂; basic: MgO, NaOH, FeO; amphoteric: Al₂O₃
- $\text{CaO} + \text{SO}_2 \rightarrow \text{CaSO}_3$; $2\text{NaOH} + \text{SiO}_2 \rightarrow \text{Na}_2\text{SiO}_3 + \text{H}_2\text{O}$; $\text{CO}_2 + \text{MgO} \rightarrow \text{MgCO}_3$

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

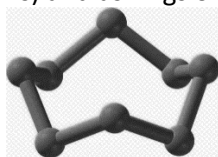
f) Sulphur and its Compounds

**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
-
- SO_2 ; it is acidic – $\text{SO}_2 + 2\text{OH}^- \rightarrow \text{SO}_3^{2-} + \text{H}_2\text{O}$
 - It will turn blue litmus paper red and will turn acidified dichromate paper from orange to green
 - H_2SO_4 ; (i) $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$, acid-base or neutralisation; (ii) $\text{H}_2\text{SO}_4 + 2\text{NH}_3 \rightarrow (\text{NH}_4)_2\text{SO}_4$, acid-base or neutralisation; (iii) $\text{H}_2\text{SO}_4 + 2\text{NaCl} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{HCl}$, acid-base or acid-salt; (iv) $2\text{H}_2\text{SO}_4 + 2\text{NaBr} \rightarrow \text{Na}_2\text{SO}_4 + \text{Br}_2 + \text{SO}_2 + 2\text{H}_2\text{O}$, redox; (v) $2\text{H}_2\text{SO}_4 + 2\text{NaI} \rightarrow \text{Na}_2\text{SO}_4 + \text{I}_2 + \text{SO}_2 + 2\text{H}_2\text{O}$, redox; $5\text{H}_2\text{SO}_4(\text{l}) + 8\text{NaI}(\text{s}) \rightarrow \text{H}_2\text{S}(\text{g}) + 4\text{I}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l}) + 4\text{K}_2\text{SO}_4(\text{s})$, redox
 - Na_2S , S^{2-}
 - Add dilute HCl and then BaCl_2 (aq), white precipitate seen

(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



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- 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: $\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$; 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:

$$\text{Zn} + \text{S} \rightarrow \text{ZnS}$$

$$\text{Cu} + \text{S} \rightarrow \text{CuS}$$

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

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

- **Sulphur dioxide (SO₂)** is a colourless gas with a simple molecular structure; like most oxides of non-metals, it is acidic and it will turn damp blue litmus paper red; it dissolves in water to make the weak acid **H₂SO₃ (sulphuric (IV) acid)**:
$$\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{SO}_3(\text{aq}) \rightleftharpoons 2\text{H}^+(\text{aq}) + \text{SO}_3^{2-}(\text{aq})$$
 - This can be used as a test for sulphur dioxide although other acidic gases, such as HCl, also do this (see Unit 5 – Acids, Bases and Salts)
 - SO₂ also reacts with alkalis to make salts containing the **sulphate (IV) ion (SO₃²⁻)**:
$$\text{SO}_2 + 2\text{OH}^- \rightarrow \text{SO}_3^{2-}$$
- **Sulphate (IV) ions (SO₃²⁻)** react with acids to give off sulphur dioxide gas
 - this can be used a test for sulphate (IV) ions (see Unit 5 – acids, bases and salts); $\text{SO}_3^{2-} + 2\text{H}^+ \rightarrow \text{SO}_2 + \text{H}_2\text{O}$
 - this reaction can also be used in the laboratory reparation of sulphur dioxide
- Sulphate (IV) ions (SO₃²⁻), sulphuric (IV) acid and sulphur dioxide (SO₂) all contain oxygen in the +4 oxidation state; sulphur is more stable in the +6 oxidation state (in the form of SO₃, sulphate ions (SO₄²⁻) and H₂SO₄); 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 SO₂ 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)
- SO₂ can be oxidised to SO₃ in a very important industrial process called the **Contact Process**:
$$\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g});$$
 - this is an exothermic and reversible reaction; according to Le Chatelier's principle, the best yield of SO₃ will be obtained at high pressure (because the reaction involves a reduction in the number of gas particles) and low temperature (because the reaction is exothermic) (see Unit 4 – Introduction to Physical Chemistry)
 - in fact a relatively low pressure of 2 atmospheres (200 kPa) is used, because a reasonably good yield is obtained at this pressure
 - a very low temperature will cause the reaction to be very slow (see Unit 4 – Introduction to Physical Chemistry); in order to balance the need for a high temperature (to create a fast reaction) and a low temperature (to create a high yield), a compromise temperature of 450 °C is used
 - a V₂O₅ 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 (SO₃)** is a colourless gas with a simple molecular structure
 - like most oxides of non-metals, it is acidic; it dissolves readily and vigorously in water to make **sulphuric acid (H₂SO₄)**: $\text{SO}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_4(\text{aq})$
 - SO₃ is formed in the atmosphere when SO₂ in the air reacts with oxygen; acid rain is formed when SO₃ in the atmosphere dissolves in rainwater

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

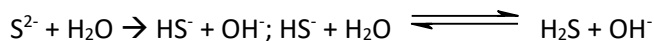
- **Sulphuric acid (H₂SO₄)** is the acid produced when SO₃ is dissolved in water; it is produced on a huge scale industrially:
 - Sulphur is oxidised to sulphur dioxide (S + O₂ → SO₂); Sulphur dioxide is oxidised to sulphur trioxide in the Contact Process: SO₂(g) + O₂(g) \rightleftharpoons 2SO₃(g); Sulphur trioxide is dissolved (indirectly) in water to make sulphuric acid
 - Sulphuric acid is a strong acid; it reacts with bases to make **sulphate salts (SO₄²⁻)** (see Topic 5 – Acids, Bases and Salts); the reaction with ammonia is carried out on an industrial scale to make ammonium sulphate, which is an important fertiliser (as a source of nitrogen for plants):
2NH₃(g) + H₂SO₄(aq) → (NH₄)₂SO₄(aq)
 - Concentrated sulphuric acid also reacts with some salts to make another acid and another salt (see Unit 5 – Acids, Bases and Salts); this is used in the laboratory preparations of HCl and HNO₃ and in the industrial preparation of phosphoric acid:
H₂SO₄ + 2KCl → K₂SO₄ + 2HCl
H₂SO₄ + 2KNO₃ → K₂SO₄ + 2HNO₃
3H₂SO₄ + Ca₃(PO₄)₂ + → 3CaSO₄ + 2H₃PO₄
 - Concentrated sulphuric acid is also an oxidising agent; it can oxidise bromide ions to bromine and iodide ions to iodine:
H₂SO₄ + 2Br⁻ + 2H⁺ → SO₂ + Br₂ + 2H₂O (S reduced from +6 to +4; Br oxidised from -1 to 0)
H₂SO₄ + 8I⁻ + 8H⁺ → H₂S + 4I₂ + 4H₂O (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 + 2H₂SO₄ → CuSO₄ + SO₂ + 2H₂O (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 (SO₄²⁻)** are formed when sulphuric acid reacts with bases; they are also formed when SO₂ or SO₃ ions are oxidised
 - the presence of SO₄²⁻ ions in solution can be easily detected by adding dilute hydrochloric acid followed by aqueous barium chloride; a white precipitate of barium sulphate is formed (see Unit 8 – Solubility and Precipitation Reactions): Ba²⁺(aq) + SO₄²⁻(aq) → BaSO₄(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 (S²⁻)

- Sulphides are formed when sulphur reacts with metals; sulphides are commonly found in minerals in the earth's crust; the most abundant are ZnS, CuS and PbS; sulphides contain the sulphide ion S²⁻

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

- 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 ($(\text{NH}_4)_2\text{S}$); the sulphide ion is strongly basic and reacts with water:



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 $\text{CuS}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CuCl}_2(\text{aq}) + \text{H}_2\text{S}(\text{g})$

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



Test your knowledge 7.2 – understanding sulphur and its compounds

- Explain why SO_2 and SO_3^{2-} ions are reducing agents and state a use for these species which relies on this property
 - State and justify the conditions used in the conversion of SO_2 to SO_3 in the Contact process
 - State the acid-base properties of sulphur trioxide and explain how acid rain is produced
 - Give a use for sulphuric acid (i) as an acid; (ii) as an oxidising agent; (iii) as a dehydrating agent
 - Write an equation to show how zinc reacts with sulphur
 - Write an equation to show how zinc sulphide reacts with dilute hydrochloric acid
 - Write an equation to show how zinc sulphide reacts with oxygen and explain why this reaction is useful
 - State how the oxidation number of S changes during the reactions above
 - State how hydrogen sulphide gas can be identified
- The S can be oxidised from +4 to +6; SO_2 is added to wine and sulphate (IV) ions are added to food to prevent oxidation
 - Low temperature gives higher yield of SO_3 but slows down reaction so compromise temperature of 450 oC used; high pressure gives higher yield of SO_3 and fast reaction but is expensive so 200 kPa used; V_2O_5 catalyst speeds up reaction
 - It is an acid and reacts with water in the air to form acid rain: $\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$
 - (i) it reacts ammonia to make ammonium sulphate which is a fertiliser and it can be used in the laboratory preparation of HCl and HNO_3 ; (ii) it is used in the laboratory preparation of SO_2 ; (iii) it is used to dry the SO_2 produced in the laboratory preparation of SO_2
 - $\text{Zn} + \text{S} \rightarrow \text{ZnS}$ (S reduced from 0 to -2)
 - $\text{ZnS} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2\text{S}$ (not a redox reaction)
 - $2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2$ (S oxidised from -2 to +4) this is the first stage in the extraction of Zn from its ore
 - See above
 - It smells of rotten eggs

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?**

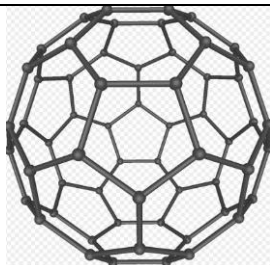
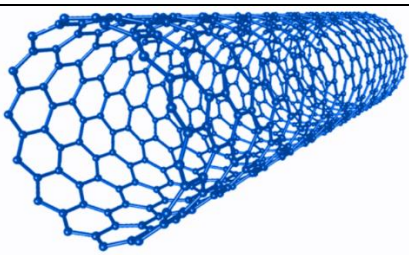
- Describe the structures of diamond and graphite; explain how their different structures result in diamond and graphite having different properties
 - Describe how carbon dioxide reacts with water and alkalis
 - Describe a test for carbon dioxide gas
 - Describe a test for carbonate ions in solution
- Diamond has a giant covalent structure; each C atom attached to four other C atoms in a tetrahedral shape; graphite has a giant covalent structure; each C atom attached to three other C atoms in a trigonal planar shape; this results in hexagonal layers; the layers are held together by Van der Waal's forces; each C has one extra electron in its outer shell which is delocalised; graphite conducts electricity due to its delocalised electrons but diamond does not; graphite is soft as the layers can slide over each other but diamond is hard because all atoms are fixed in place
 - $\text{CO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$; $\text{CO}_2 + 2\text{NaOH} \rightarrow \text{Na}_2\text{CO}_3 + \text{H}_2\text{O}$ (CO_2 is acidic)
 - Turns limewater milky and then colourless again
 - Bubbles on addition of HCl; CO_2 given off; precipitate formed on addition of BaCl_2 but precipitate dissolves in HCl

(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**

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

- 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 C₆₀
 - or as **nanotubes**, which are long cylinders of carbon atoms

Name	Buckminster-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-bituminous coal** (71 – 77% carbon) and **bituminous coal** (77 – 87% carbon)
 - the coal eventually being transformed into a metamorphic rock called **anthracite** (more than 87% carbon)
 - these processes take place over millions of years
- 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 gasification (heating at 700 °C in a limited supply of oxygen) ($C + H_2O \rightarrow CO + H_2$); synthesis gas is itself a fuel and can also be used in the manufacture of methanol and hydrogen)

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

(ii) Carbon dioxide (CO₂), carbonate ions (CO₃²⁻) and hydrogencarbonate ions (HCO₃⁻)

- **Carbon dioxide (CO₂)** 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)
$$\text{CO}_2 + \text{OH}^- \rightarrow \text{HCO}_3^-; \text{CO}_2 + 2\text{OH}^- \rightarrow \text{CO}_3^{2-} + \text{H}_2\text{O}$$

Eg $\text{CO}_2 + \text{NaOH} \rightarrow \text{NaHCO}_3; \text{CO}_2 + 2\text{NaOH} \rightarrow \text{Na}_2\text{CO}_3 + \text{H}_2\text{O}$
This reaction is the basis for the test for carbon dioxide: CO₂ 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):
$$\text{CO}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$$

$$\text{CaCO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{HCO}_3)_2(\text{aq})$$
 - 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 (CO₃²⁻)** are commonly found in rocks
 - carbonate ions are formed when metal oxides (which are basic) react with CO₂ (which is acidic)
eg $\text{MgO} + \text{CO}_2 \rightarrow \text{MgCO}_3$
 - carbonate ions can also be formed in solution by dissolving carbon dioxide in alkalis
$$\text{CO}_2 + 2\text{OH}^- \rightarrow \text{CO}_3^{2-} + \text{H}_2\text{O}$$
 - carbonate ions react with acids to form carbon dioxide: $\text{CO}_3^{2-} + 2\text{H}^+ \rightarrow \text{CO}_2 + \text{H}_2\text{O}$ (see Unit 5 – Acids, Bases and Salts); this is commonly used in the preparation of carbon dioxide
 - **Sodium carbonate (Na₂CO₃)** is a weak base, due to the CO₃²⁻ 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 CO₃²⁻ in aqueous solution; it is used in the manufacture of glass (by mixing it with silicon dioxide); it is also used as a water softener and to raise the pH of water; it is also very useful for making standard solutions because it is stable in its most common form Na₂CO₃·10H₂O
 - **Calcium carbonate (CaCO₃)** is a weak base due to the CO₃²⁻ ion; it is insoluble in water and is the main component of limestone, chalk and marble; it is added to soil to reduce its acidity and is widely used in building materials, although it is damaged by acids; it is also used in cement; it decomposes on heating into CaO; $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$; it is therefore used to make CaO and is generally added during the manufacture of steel to produce the CaO
- **Hydrogencarbonate ions (HCO₃⁻)** are bases made by the partial neutralisation of CO₂ with alkalis or by the partial neutralisation of carbonate ions with acids:
$$\text{CO}_2 + \text{OH}^- \rightarrow \text{HCO}_3^-$$

$$\text{CO}_3^{2-} + \text{H}^+ \rightarrow \text{HCO}_3^-$$
 - HCO₃⁻ ions react with bases to make carbonates and with acids to make CO₂; they are therefore amphoteric (see Unit 5 – Acids, Bases and Salts); HCO₃⁻ ions also react with themselves to make carbonates and CO₂:
$$\text{HCO}_3^- + \text{OH}^- \rightarrow \text{CO}_3^{2-} + \text{H}_2\text{O}$$

$$\text{HCO}_3^- + \text{H}^+ \rightarrow \text{CO}_2 + \text{H}_2\text{O};$$


$$2\text{HCO}_3^- \rightarrow \text{CO}_2 + \text{CO}_3^{2-} + \text{H}_2\text{O}$$

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

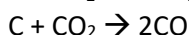
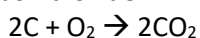
- **Sodium hydrogencarbonate (NaHCO₃)** is a weak base due to the HCO₃⁻ ion; it is also known as baking soda or baking powder; NaHCO₃ reacts with itself on heating to release CO₂: $2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$; when added to dough, the CO₂ is slowly released into the dough on heating, causing it to rise; it is therefore very useful in baking; in water it behaves predominantly as a mild base and is taken in tablets to reduce stomach acidity (ie treat indigestion); it is also used to raise the pH of swimming pools

(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

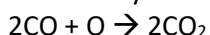
electrons: 

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



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

- it reacts readily with oxygen to produce CO₂:



Test your knowledge 8.2 – understanding carbon and its compounds

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

- Peat is plant and animal matter partially decayed in acidic, anaerobic conditions (<60% C); lignite is a sedimentary rock formed when peat is compressed (60 – 70% C); continued pressure converts lignite into coal (70 – 87% C) and continued high temperature and pressure converts coal into a metamorphic rock called anthracite (>87% C)
- Coke is produced by the destructive distillation of coal; coal gas, coal tar and coal oil are also produced
- Mixture of CO and H₂, produced by heating coke at 700 oC in a limited supply of oxygen
- $\text{CO}_2 + 2\text{OH}^- \rightarrow \text{CO}_3^{2-} + \text{H}_2\text{O}$; $\text{CO}_3^{2-} + 2\text{H}^+ \rightarrow \text{CO}_2 + \text{H}_2\text{O}$; $2\text{HCO}_3^- \rightarrow \text{CO}_3^{2-} + \text{CO}_2 + \text{H}_2\text{O}$ or $\text{HCO}_3^- + \text{H}^+ \rightarrow \text{CO}_2 + \text{H}_2\text{O}$ and $\text{HCO}_3^- + \text{OH}^- \rightarrow \text{CO}_3^{2-}$
- Add dilute HCl to CaCO₃; collect CO₂ by downward delivery
- It is not flammable and more dense than air
- Na₂CO₃ – water softener and used to make glass; CaCO₃ – used in steel manufacture and to neutralise acidic soil; NaHCO₃ – baking powder

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

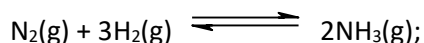
h) Nitrogen and its Compounds

**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
- It is simple molecular, forming diatomic molecules (N₂) in which two N atoms are joined by triple covalent bonds; the molecules are held together by weak Van der Waal's forces
 - It is simple molecular, consisting of (NH₃) molecules in which the N atom is attached to three H atoms by covalent bonds; the molecules are held together by hydrogen bonds
 - NH₃ + HCl → NH₄Cl
 - It turns red litmus blue and forms a white smoke in the presence of concentrated HCl
 - If warmed with NaOH, it gives off NH₃ which has a pungent smell

(i) Nitrogen (N₂)

- Nitrogen has a simple molecular structure (see Unit 2 – Particles, Bonding and Structure); it exists as a diatomic gas (N₂) and makes up 80% of the atmosphere; the bond between the nitrogen atoms is a triple covalent bond which is very strong and hard to break; this makes nitrogen a very unreactive gas; it is non-polar and so not very soluble in water
- Industrially, pure nitrogen is prepared by the fractional distillation of liquid air in the same process as that used to prepare oxygen industrially (see above)
- The main use of nitrogen is to manufacture **ammonia (NH₃)** by the **Haber process**:



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

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

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

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

NH₄⁺ ions are a very important source of nitrogen, which plants need; bacteria in the soil gradually oxidise NH₄⁺ ions to NO₃⁻ ions, which plants can absorb through their roots
- Ammonia is also a component of cleaners; alkaline solutions are useful for breaking down fats into water soluble products
- The **ammonium ion (NH₄⁺)** is acidic and ammonium salts (as salts of weak bases) are therefore acidic; NH₄⁺ ions can react with bases to form NH₃ gas; this is used as a test for NH₄⁺ 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 NH₃**; solid Ca(OH)₂ is heated with solid NH₄Cl and the following reaction takes place: $\text{Ca}(\text{OH})_2(\text{s}) + 2\text{NH}_4\text{Cl}(\text{s}) \rightarrow \text{CaCl}_2(\text{s}) + 2\text{NH}_3(\text{g}) + 2\text{H}_2\text{O}(\text{l})$
- In NH₃ and NH₄⁺, 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; NH₄⁺ ions are thus mild reducing agents; one oxidising agent which can be used for this purpose is NaNO₂; NO₂⁻ ions are known as nitrite (III) ions:
$$\text{NaNO}_2(\text{s}) + \text{NH}_4\text{Cl}(\text{s}) \rightarrow \text{N}_2(\text{g}) + \text{NaCl}(\text{s}) + 2\text{H}_2\text{O}(\text{l})$$

The N in NH₄⁺ is oxidised from -3 to 0; the N in NO₂⁻ is reduced from +3 to 0; N₂ 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

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

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

- (a) The air is cooled to a very low temperature to ensure that N_2 , O_2 and Ar all condense; liquid air run into base of fractionating column; temperature allowed to rise slowly; Ar boils first, then N_2 , so N_2 collected at top of column
- (b) NH_4Cl reacts with $NaNO_2$; redox reaction; N_2 not very soluble in water so can be collected over water
- (c) $Ca(OH)_2$ reacts with NH_4Cl ; acid-base reaction; NH_3 very soluble in water so not collected over water; less dense than air so collected by upward delivery
- (d) Low temperature gives higher yield of ammonia but slows down reaction so compromise temperature of $450\text{ }^\circ\text{C}$ used; high pressure gives higher yield of ammonia and fast reaction so 25 MPa used; Fe catalyst speeds up reaction
- (e) To make fertilisers

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?

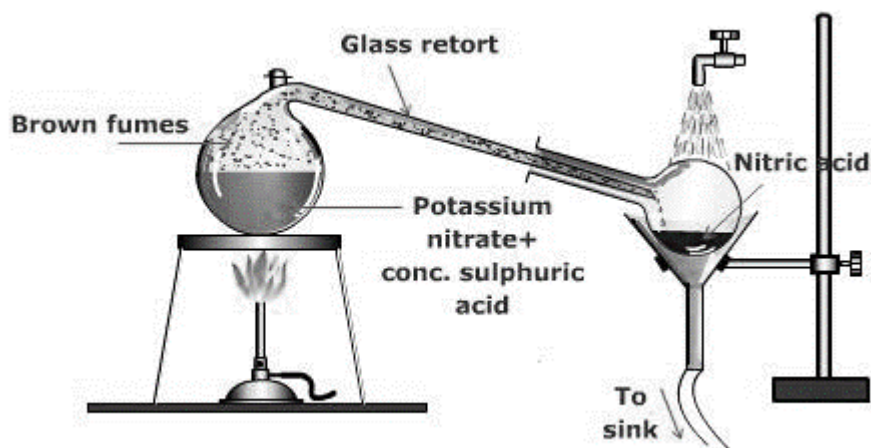
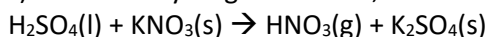
- Give the formula of nitric acid and the nitrate ion
 - Discuss the acid-base properties of nitric acid and nitrate ions
 - Write an equation to show how nitric acid reacts with ammonia
 - Describe a simple chemical test for nitrate ions
-
- HNO_3 , NO_3^-
 - Nitric acid is a strong acid, so nitrate ions are neutral
 - $HNO_3 + NH_3 \rightarrow NH_4NO_3$
 - Add aluminium powder, NaOH and heat; a pungent gas (NH_3) is given off which turns damp red litmus paper blue

(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 $83\text{ }^\circ\text{C}$; 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: $HNO_3 + NH_3 \rightarrow NH_4NO_3$

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

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



An animation of this process can be seen on www.youtube.com/watch?v=OxZZA_BbmNs
Gentle heating of the mixture causes the HNO_3 to evaporate; it is distilled off and can be condensed separately

- Nitrate (V) ions (NO_3^-)** are found in dilute nitric acid and in all salts of nitric acid; in both nitric acid and nitrate (V) ions, the oxidation number of the N is +5; nitrate (V) ions should be distinguished from nitrate (III) ions (NO_2^-) in which the N has an oxidation number of +3; however NO_3^- ions are commonly known simply as “nitrate ions” because they are much more common and stable than NO_2^- ions, which are always referred to as nitrate (III) ions
- All nitrates are soluble in water so do not give precipitation reactions (see Unit 8 – Solubility and Precipitation Reactions); they are salts of a strong acid so are not bases (see Unit 5 – Acids, Bases and Salts); nitrate ions can behave as oxidising agents; when heated with Al and NaOH the NO_3^- ion can be reduced to NH_3 , which is given off as a pungent gas which turns damp red litmus paper blue; the N is reduced from +5 (in the NO_3^-) to -3 (in the NH_3); this is the standard test for NO_3^- ions (see Unit 6 – Redox Reactions)
- Nitrates are unstable on heating – they undergo **thermal decomposition**; the N is reduced and the O is oxidised:
 - $NaNO_3$ and KNO_3 decompose to give nitrate (III) salts and oxygen gas:
$$2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$$

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

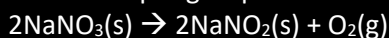
Write an equation for the reactions occurring when both salts are heated; in each case, state what is oxidised and what is reduced

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

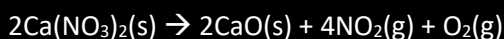
What is responsible for the brown colour in the second reaction?

Chemicals needed per group: calcium nitrate and sodium nitrate (around 2 g per group), each container with its own spatula

Apparatus needed per group: 2 test tubes, 1 tongs, 1 Bunsen burner, 1 splint



The N is reduced from +5 to +3 and the O is oxidised from -2 to 0



The N is reduced from +5 to +4 and the O is oxidised from -2 to 0

The glowing splint re-lights, confirming the presence of oxygen

The brown gas is NO_2

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

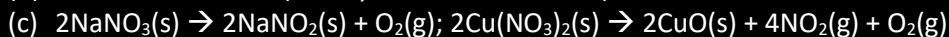


Test your knowledge 10.3 – Nitric acid and nitrates

- (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

(a) By adding concentrated sulphuric acid to potassium nitrate: $\text{H}_2\text{SO}_4(\text{l}) + \text{KNO}_3(\text{s}) \rightarrow \text{HNO}_3(\text{g}) + \text{K}_2\text{SO}_4(\text{s})$; acid-base or acid-salt reaction

(b) To make fertilisers (mainly ammonium nitrate)



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 (H₂), oxygen (O₂), nitrogen (N₂), chlorine (Cl₂), hydrogen chloride (HCl), sulphur dioxide (SO₂), carbon dioxide (CO₂) and ammonia (NH₃); 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
- H₂: $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$ (redox or metal-acid); O₂: $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$ (redox or disproportionation); N₂: $\text{NH}_4\text{Cl} + \text{NaNO}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O} + \text{NaCl}$ (redox); Cl₂: $\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + 2\text{H}_2\text{O}$ (redox); HCl: $\text{H}_2\text{SO}_4 + 2\text{KCl} \rightarrow \text{K}_2\text{SO}_4 + 2\text{HCl}$ (acid-salt); SO₂: either $\text{Na}_2\text{SO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{SO}_2 + \text{H}_2\text{O}$ (acid-base) or $\text{Cu} + 2\text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{SO}_2 + 2\text{H}_2\text{O}$ (redox); CO₂: $\text{CaCO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}$ (acid-base); NH₃: $\text{Ca(OH)}_2 + 2\text{NH}_4\text{Cl} \rightarrow \text{CaCl}_2 + 2\text{NH}_3 + 2\text{H}_2\text{O}$ (acid-base)
- H₂ burns with a squeaky pop; O₂ relights a glowing splint; there is no specific test for N₂; Cl₂ turns blue litmus red and then white; HCl will give white fumes in contact with a strip of filter paper soaked in NH₃; SO₂ will turn dichromate paper from orange to green; CO₂ will turn limewater milky and then colourless again; NH₃ will give white fumes in contact with a strip of filter paper soaked in HCl

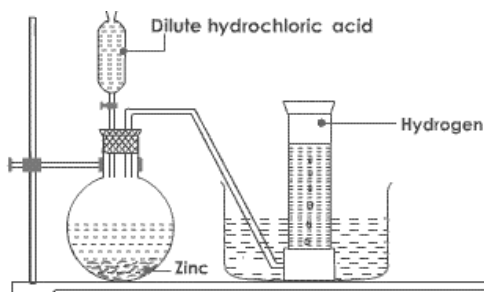
(i) General Principles of Collection of Gases

- Earlier in this topic you have studied the chemical reactions used to prepare the following gases in the laboratory: hydrogen (H₂), oxygen (O₂), nitrogen (N₂), chlorine (Cl₂), hydrogen chloride (HCl), sulphur dioxide (SO₂), carbon dioxide (CO₂) and ammonia (NH₃)
- 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 NH₃) 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

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

(ii) hydrogen

- **Hydrogen (H₂)** 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): $\text{Zn} + 2\text{HCl} \rightarrow \text{Zn} + \text{ZnCl}_2 + \text{H}_2$ (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:



www.blogspot.com



Demonstration 11.2: Prepare H₂ 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 mol dm⁻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 H₂ 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 H₂ 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?

Equipment and chemicals needed as described 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, a gas jar with lid, clamp, stand, boss, 2 mol dm⁻³ HCl, zinc granules)

Bubbles can be clearly seen when the HCl is added, and bubbles should be seen entering the gas jar; the water level in the gas jar should drop

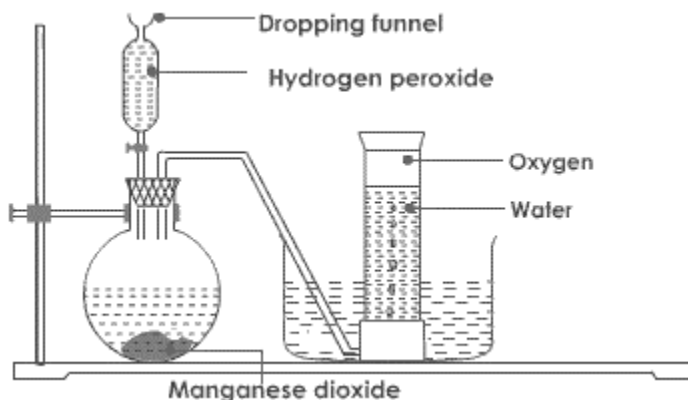
- $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$
- Zn is oxidised (from 0 to +2) and the H⁺ in the acid is reduced from +1 to 0
- it is insoluble in water because it is non-polar
- hydrogen gas is less dense than air so it rises, if you remove the lid from underneath the jar, most of the gas will remain in the jar; if you remove the lid from the top of the jar, the gas will all escape
- It burns with a squeaky pop
- The hydrogen is reacting with oxygen: $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$

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

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

(iii) oxygen

- **Oxygen (O₂)** is prepared in the laboratory by the decomposition of hydrogen peroxide in the presence of a catalyst (MnO₂)
 - it is a redox reaction (disproportionation): $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$ (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:



Demonstration 11.3: Prepare O₂ gas in the laboratory

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

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

Write an equation for the reaction taking place

What is oxidised and what is reduced?

Why can O₂ be collected over water?

What is the test for oxygen gas?

What reaction is taking place during this test?

Equipment and chemicals needed as described above (wear gloves for this experiment) (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, a gas jar with lid, clamp, stand, boss, 2 moldm⁻³ H₂O₂, MnO₂ powder)

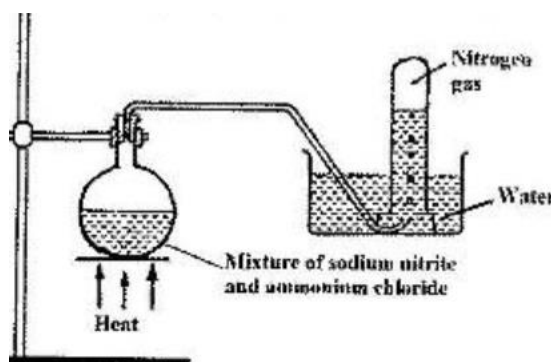
Bubbles can be clearly seen when the H₂O₂ is added, and bubbles should be seen entering the gas jar; the water level in the gas jar should drop

- $2\text{H}_2\text{O}_2(\text{l}) \rightarrow 2\text{H}_2\text{O} + \text{O}_2(\text{g})$
- In this reaction the O is both oxidised (from -1 to 0, in O₂) and reduced (from -1 to -2, in H₂O)
- it is insoluble in water because it is non-polar
- It relights a glowing splint
- The wood in the splint is reacting with the oxygen in a combustion reaction

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

(iv) Nitrogen

- **Nitrogen (N₂)** is prepared in the laboratory by the reaction of ammonium chloride with sodium nitrate (III)
 - it is a redox reaction: $\text{NH}_4\text{Cl} + \text{NaNO}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O} + \text{NaCl}$ (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:



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Demonstration 11.4: Prepare N₂ 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 NaNO₂ and 11 g of NH₄Cl in around 200 cm³ 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 N₂ be collected over water?

Equipment and chemicals needed as described above (a flask, a dropping funnel, a delivery tube, a bung to fit the flask with a delivery tube, Bunsen, tripod, gauze, a trough of water, a gas jar with lid, clamp, stand, boss, 14 g NaNO₂, 11 g of NH₄Cl)

Bubbles should be visible shortly after heating starts; bubbles should be seen entering the gas jar; the water level in the gas jar should drop

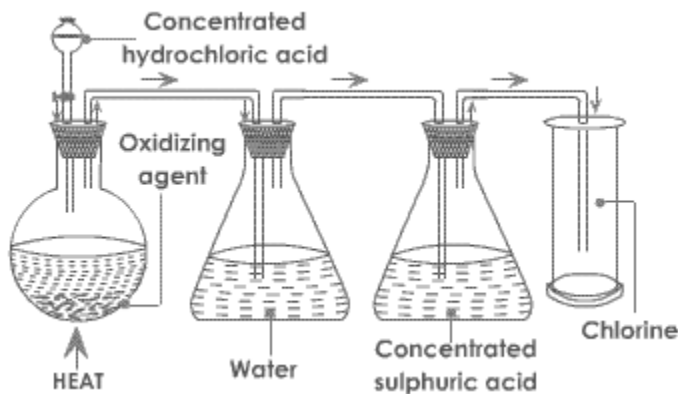
- $\text{NH}_4\text{Cl} + \text{NaNO}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O} + \text{NaCl}$
- N in NH_4^+ is oxidised from -3 to 0 and N in NO_2^- is reduced from +3 to 0
- it is insoluble in water because it is non-polar

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

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

(v) chlorine

- **Chlorine (Cl₂)** is prepared in the laboratory by the reaction of MnO₂ with concentrated HCl
 - this is a redox reaction: $\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + 2\text{H}_2\text{O}$ (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 KMnO₄ can be used as an alternative to MnO₂
 - chlorine is denser than air ($m_r = 71$) so should be collected by downward delivery:



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Demonstration 11.5: Prepare Cl₂ gas in the laboratory

(caution – concentrated HCl and concentrated H₂SO₄ 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, three delivery tubes, a gas jar, Bunsen burner, tripod, gauze)
- 2) Fill the dropping funnel with concentrated HCl; add around 5 g of MnO₂ into the round-bottomed flask; add around 50 cm³ of water to the first conical flask and 50 cm³ of concentrated sulphuric acid to the second conical flask
- 3) Gradually add the concentrated HCl to the MnO₂ 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 Cl₂ 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 Cl₂ bubbled through first water and then concentrated H₂SO₄?

Why is the gas then collected using downward delivery?

Why does Cl₂ turn damp litmus paper red and then bleach it?

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

Equipment and chemicals needed: a round-bottomed flask, two conical flasks, a dropping funnel, three bungs to fit the flasks, each with two holes, three delivery tubes, a gas jar, Bunsen burner, tripod, gauze, MnO₂, conc HCl, conc H₂SO₄, damp blue litmus paper) **you must use gloves for this experiment**

Bubbles can be clearly seen when the conc HCl is added; heating may not be necessary; bubbles should be seen passing through the water and the sulphuric acid; the gas may appear as a pale green colour in the gas jar

- $\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + 2\text{H}_2\text{O}$
- The Mn is reduced from +4 to +2; the Cl⁻ is oxidised to Cl₂ (-1 to 0)
- Water removes the HCl and concentrated H₂SO₄ removes the water
- Chlorine gas is denser than air
- $\text{Cl}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCl}(\text{aq}) + \text{HClO}(\text{aq})$; HCl contains H⁺ and turns litmus red; HClO is a bleaching agent and turns it white

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)

- Write an equation for the reactions used to prepare the following gases in the laboratory:
 - hydrogen
 - oxygen
 - nitrogen
 - chlorine
- For each of the gases above, state how the gas is collected and explain any differences in the way the gases are collected
- Describe a simple chemical test which could confirm the identify of:
 - hydrogen
 - oxygen
 - chlorine

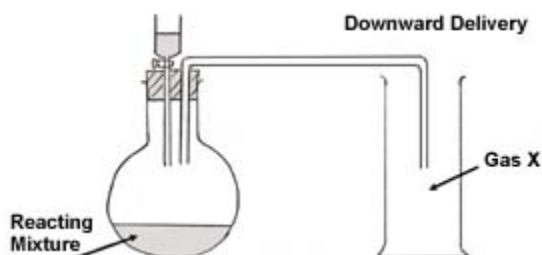
- (i) H_2 : $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$ (redox or metal-acid); (ii) O_2 : $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$ (redox or disproportionation); (iii) N_2 : $\text{NH}_4\text{Cl} + \text{NaNO}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O} + \text{NaCl}$ (redox); (iv) Cl_2 : $\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + 2\text{H}_2\text{O}$ (redox)
- H_2 , O_2 and N_2 have limited solubility in water so are collected over water; Cl_2 reacts with water and is denser than air so is collected by downward delivery
- H_2 burns with a squeaky pop; O_2 relights a glowing splint; Cl_2 turns blue litmus red and then white

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

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: $\text{H}_2\text{SO}_4 + 2\text{KCl} \rightarrow \text{K}_2\text{SO}_4 + 2\text{HCl}$
 - 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 gas jar and place a lid on the gas jar
- 5) Confirm that the gas is HCl by dipping briefly some damp blue litmus paper into the gas jar, and then a piece of filter paper soaked in 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

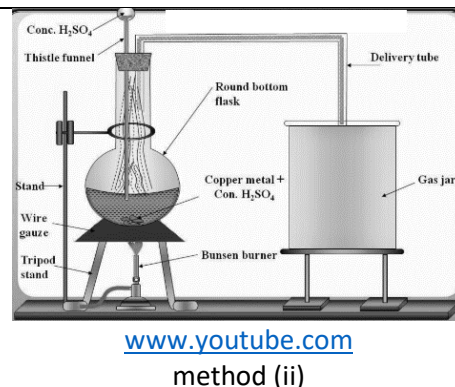
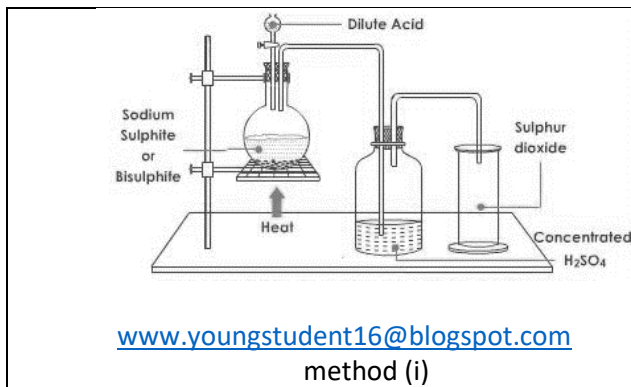
Equipment and chemicals needed: round-bottomed flask, dropping funnel, bung for flask with 2 holes, delivery tube, gas jar, spatula, dropping pipette, gas jar, clamp, stand and boss; solid KCl, conc H_2SO_4 , filter paper soaked in conc NH_3 (you must use gloves for this experiment and use a fume cupboard if you have one)

- HCl collected by downward delivery as it is denser than air
- HCl not collected over water as it is very soluble in water
- $\text{H}_2\text{SO}_4(\text{l}) + 2\text{KCl}(\text{s}) \rightarrow \text{K}_2\text{SO}_4(\text{s}) + 2\text{HCl}(\text{g})$
- $\text{HCl}(\text{g}) \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ (H^+ turns the blue litmus red)
- $\text{NH}_3(\text{g}) + \text{HCl}(\text{g}) \rightarrow \text{NH}_4\text{Cl}(\text{s})$ (NH_4Cl is the white smoke)

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

(vii) Sulphur dioxide

- **Sulphur dioxide (SO₂)** can be prepared in the laboratory by in one of two ways:
 - (i) by the addition of dilute HCl to a solid sample of Na₂SO₃; this is an acid-base reaction:
$$\text{Na}_2\text{SO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{SO}_2 + \text{H}_2\text{O}$$
 - (ii) by the reduction of concentrated sulphuric acid using copper metal; this is a redox reaction:
$$\text{Cu} + 2\text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{SO}_2 + 2\text{H}_2\text{O}$$
 (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 H₂SO₄ in order to remove any water vapour
 - sulphur dioxide is denser than air (m_r = 64) so it should be collected by downward delivery:



Demonstration 12.2: Prepare SO₂ 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 Na₂SO₃ into the round-bottomed flask; add around 50 cm³ of concentrated sulphuric acid to the second conical flask
- 3) Gradually add the dilute HCl, and start heating the mixture; you will observe a gas being given off which will eventually collect in the gas jar
- 4) When a reasonable quantity of gas has been collected in the gas jar, close the tap of the dropping funnel, remove the delivery tube from the gas jar and place a lid on the gas jar
- 5) Confirm that the gas is SO₂ by dipping briefly some damp blue litmus paper, and then some acidified dichromate paper, into the gas jar

Write an equation for the reaction taking place

What type of reaction is this?

Why is the gas bubbled through concentrated H₂SO₄?

Why is the gas collected by downward delivery?

Why is the gas not collected over water?

State and explain the effect of SO₂ on damp blue litmus paper and on acidified dichromate paper

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

Equipment and chemicals needed: (round-bottomed flask, dropping funnel, conical flasks, two delivery tubes, two bungs to fit the flasks, each with two holes, a gas jar, clamp, stand, boss, Bunsen, tripod, gauze, 5 g Na₂SO₃, 50 cm³ 1 mol dm⁻³ HCl, blue litmus paper, dichromate paper); **you must use gloves for this experiment**

Bubbles can be clearly seen when the dilute HCl is added; heating should increase the rate of production of bubbles and bubbles should be seen passing through the sulphuric acid

- $\text{Na}_2\text{SO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
- Acid-base reaction
- concentrated H₂SO₄ removes the water
- SO₂ is denser than air
- SO cannot be collected over water as it is soluble in water
- $\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{SO}_3(\text{aq}) \rightleftharpoons 2\text{H}^+(\text{aq}) + \text{SO}_3^{2-}(\text{aq})$; H⁺ and turns litmus red
- SO₂ is a reducing agent; it gets oxidised to SO₄²⁻ ions and reduces dichromate ions to Cr³⁺

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



Demonstration 12.3: Prepare SO₂ gas in the laboratory (Part 2)

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

- 1) Set up the apparatus as shown in the diagram above (a round-bottomed flask, a 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 H₂SO₄; 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 SO₂ by dipping briefly some damp blue litmus paper, and then some acidified dichromate paper, into the gas jar

Write an equation for the reaction taking place

What is oxidised and what is reduced?

Equipment and chemicals (round-bottomed flask, dropping funnel, delivery tube, bung to fit the flask with two holes, gas jar, Bunsen burner with tripod and gauze, clamp, stand, boss, conc H₂SO₄ (50 cm³), Cu turnings (5 g), blue litmus paper, dichromate paper) **you must use gloves for this experiment**

Bubbles should be seen when the mixture is heated

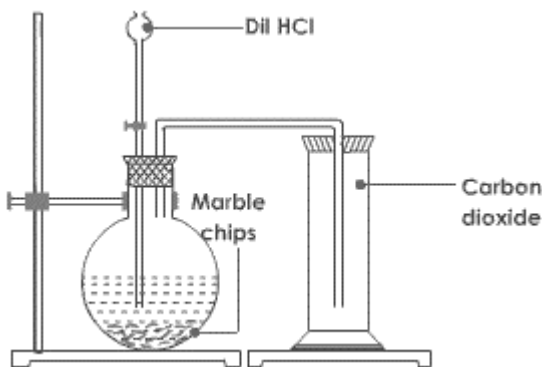
- $\text{Cu} + 2\text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{SO}_2 + 2\text{H}_2\text{O}$
- S reduced from +6 to +4; Cu oxidised from 0 to +2

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

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

(viii) Carbon dioxide

- **carbon dioxide (CO₂)** is generally prepared in the laboratory by reacting carbonates with acids
 - this is an acid-base reaction: $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{CO}_2 + \text{H}_2\text{O}$
 - 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 CO₂ gas in the laboratory

- 1) Set up the apparatus as shown in the diagram above (a round-bottomed flask, a dropping funnel, a delivery tube, a bung to fit the flask with two holes, a gas jar, 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 CO₂ have on limewater and why?

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

Equipment and chemicals needed: (round-bottomed flask, dropping funnel, delivery tube, bung to fit the flask with two holes, gas jar, test tube, clamp with stand and boss, marble chips (5 g), 2 moldm⁻³ HCl (50 cm³), limewater, 10 cm³, splint)

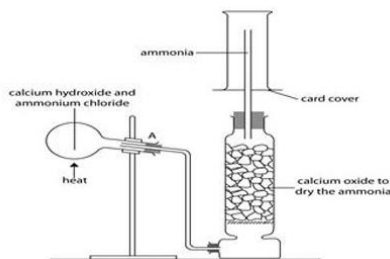
Bubbles can be clearly seen when the dilute HCl is added

- $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
- Acid-base reaction
- CO₂ is denser than air
- CO₂ turns limewater milky and then colourless again; the milky appearance is due to the insoluble calcium carbonate; $\text{CO}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$; $\text{CaCO}(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{HCO}_3)_2(\text{aq})$
- Carbon dioxide extinguishes a lit splint as it is not flammable

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

(ix) Ammonia

- **ammonia (NH₃)** 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: $\text{Ca(OH)}_2(\text{s}) + 2\text{NH}_4\text{Cl}(\text{s}) \rightarrow \text{CaCl}_2(\text{s}) + 2\text{NH}_3(\text{g}) + 2\text{H}_2\text{O}(\text{l})$
 - 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 NH₃ 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 NH₃ 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 NH₃ not be collected over water?

What is the function of the CaO?

Why must NH₃ be collected by upward delivery?

What is the test for NH₃ and how does it work?

Equipment and chemicals needed: round-bottomed flask with a bung attached to a delivery tube, CaO tower, gas jar, Bunsen burner, NH₄Cl (2 g), Ca(OH)₂ (2 g), red litmus paper

Bubbles should be visible after gently heating the mixture for a short time

- $\text{Ca(OH)}_2(\text{s}) + 2\text{NH}_4\text{Cl}(\text{s}) \rightarrow \text{CaCl}_2(\text{s}) + 2\text{NH}_3(\text{g}) + 2\text{H}_2\text{O}(\text{l})$
- A neutralisation or acid-base reaction is taking place
- Ammonia is very soluble in water so cannot be collected over water
- To remove any water from the gaseous mixture
- It is less dense than air
- It turns red litmus paper blue; NH₃ is a base and reacts with water on the paper to form OH⁻ ions

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:
- hydrogen chloride
 - sulphur dioxide
 - carbon dioxide
 - 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:
- hydrogen chloride
 - sulphur dioxide
 - carbon dioxide
 - ammonia

- (a) (i) $\text{HCl: H}_2\text{SO}_4 + 2\text{KCl} \rightarrow \text{K}_2\text{SO}_4 + 2\text{HCl}$ (acid-salt); (ii) SO_2 : either $\text{Na}_2\text{SO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{SO}_2 + \text{H}_2\text{O}$ (acid-base) or $\text{Cu} + 2\text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{SO}_2 + 2\text{H}_2\text{O}$ (redox); (iii) CO_2 : $\text{CaCO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}$ (acid-base); (iv) NH_3 : $\text{Ca(OH)}_2 + 2\text{NH}_4\text{Cl} \rightarrow \text{CaCl}_2 + 2\text{NH}_3 + 2\text{H}_2\text{O}$ (acid-base)
- (b) HCl , SO_2 , CO_2 and NH_3 are highly soluble in water so cannot be collected over water; HCl , SO_2 and CO_2 are denser than air so collected by downward delivery; NH_3 is less dense than air so collected by upward delivery
- (c) HCl will give white fumes in contact with a strip of filter paper soaked in NH_3 ; SO_2 will turn dichromate paper from orange to green; CO_2 will turn limewater milky and then colourless again; NH_3 will give white fumes in contact with a strip of filter paper soaked in HCl

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)

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

- **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
-
- (a) Mineral – substance of fixed composition found in the earth's crust; rock – mixture of minerals
 - (b) Igneous – formed by cooling lava or magma; sedimentary – formed by small rock deposits being compressed; metamorphic – formed by subjecting sedimentary rocks to high temperature and pressure
 - (c) Breaking down of rocks into smaller particles
 - (d) Hydrolysis - breaking down a rock by reacting it with water, acid or alkali; hydration - absorption of water by a rock; carbonation – dissolving a rock in carbonic acid; oxidation – reaction of a rock with oxygen

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

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

UNIT 12 – NON-METALS AND THEIR COMPOUNDS

1. Ar = simple atomic, Van der Waal's forces between atoms; Cl₂, H₂, O₂, N₂ = simple molecular, diatomic molecules with single covalent bonds (Cl-Cl and H-H) or double covalent bonds (O=O) or triple covalent bonds (N≡N between atoms; S₈ crown shaped molecule with eight S atoms held together in a ring by single covalent bonds, Van der Waal's forces between all molecules)
2. O₂: decomposition of H₂O₂ in presence of MnO₂; Cl₂: oxidation of concentrated HCl by MnO₂; N₂: oxidation of NH₄Cl by NaNO₂; redox reactions; Cl₂ collected by downward delivery as denser than air and reacts with water; N₂ and O₂ collected over water as they are not very soluble in water
3. HCl: H₂SO₄ with KCl (acid-base or acid-salt); NH₃: NH₄Cl with Ca(OH)₂ (acid-base); CO₂: CaCO₃ with HCl (acid-base); HCl and NH₃ highly soluble in water so not collected over water; HCl denser than air so collected by downward delivery; NH₃ less dense than air so collected by upward delivery; CO₂ denser than air so collected by downward delivery (or over water)
4. H₂: burns with a squeaky pop; O₂: relights a glowing splint; Cl₂: turns blue litmus paper red and then white; N₂: negative test for all other gases
5. (a) $\text{Cl}_2 + 2\text{NaOH} \rightarrow \text{NaCl} + \text{NaClO} + \text{H}_2\text{O}$
(b) $\text{Cl}_2 + 2\text{NaBr} \rightarrow 2\text{NaCl} + \text{Br}_2$
6. Fractional distillation; air is cooled until it completely liquefies and inserted into the base of a fractionating column; N₂ has a lower boiling point and is collected at the top of the fractionating column
7. $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$; 450 °C (compromise temperature: high temperature gives faster reaction but lower yield); 25 MPa (high pressure gives fast reaction and good yield); Fe catalyst to increase rate of reaction
8. (a) $2\text{Ca}(\text{NO}_3)_2 \rightarrow 2\text{CaO} + 4\text{NO}_2 + \text{O}_2$
(b) $2\text{NaNO}_3 \rightarrow 2\text{NaNO}_2 + \text{O}_2$
(c) $2\text{CuS} + 3\text{O}_2 \rightarrow 2\text{CuO} + 2\text{SO}_2$
9. Peat compressed into lignite and then into coal; over millions of years it is further compressed at high temperatures into anthracite
10. Igneous – formed from cooling lava or magma; sedimentary – formed by compression and sedimentation of small rock deposits; metamorphic – formed by further compression of sedimentary rock at high temperature
11. Breaking down of rocks into smaller pieces (hydration, hydrolysis, carbonation and oxidation)