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| UNIT 12 **NON-METALS AND THEIR COMPOUNDS**  **Answers** |

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

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

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| **Image result for test iconTest your knowledge 1.2: Group 0 – understanding the noble gases** |
| 1. Eg ionisation energies very high so cannot lose electrons, outer shell full so cannot share or accept electrons easily 2. Increase on descending the group; more electrons per atom so stronger Van der Waal’s forces between atoms 3. Increase on descending the group; mass of nucleus increases much faster than volume of atoms 4. Helium – hot air balloons/oxygen tanks for diving; neon – lighting; argon – inert atmosphere |

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| **Image result for test iconTest your knowledge 1.3: Group 7 – understanding the halogens (physical properties)** |
| 1. Eg ionisation energies very high so cannot lose electrons, outer shell full so cannot share or accept electrons easily 2. Increase on descending the group; more electrons per atom so stronger Van der Waal’s forces between atoms 3. Increase on descending the group; mass of nucleus increases much faster than volume of atoms 4. 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)?***

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| **Image result for test iconTest your knowledge 2.1: Group 7 – understanding redox properties of halogens and halides** |
| 1. 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 2. Increases down the group from fluoride to iodide; more shells means more shielding, so electrons in the outer shell as less strongly held and can be more easily lost 3. The orange colour disappears and a yellow/brown colour appears; this is because I2 is being produced; Br2 + 2I- 🡪 2Br- + I2; Br2 is a stronger oxidising agent than I2 so Br will displace I from its compounds 4. No reaction; Br2 is a weaker oxidising agent than I2 so Br will not displace Cl from its compounds |

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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 2.2: Carry out halogen displacement reactions** |
| 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)    The halogen changes during the following reactions:  Cl2 with KBr and KI, Br2 with KI:  Cl2 + 2Br- 🡪 2Cl- + Br2; Cl2 + 2I- 🡪 2Cl- + I2; Br2 + 2I- 🡪 2Br- + I2; |

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

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| Image result for test icon**Test your knowledge 3.1: Group 7 – comparing the reducing power of halides** |
| 1. H2SO4(l) + 2KCl(s) 🡪 K2SO4(s) + 2HCl(g) 2. 2H2SO4(l) + 2KBr(s) 🡪 SO2(g) + Br2(g) + 2H2O(l) + K2SO4(s) 3. 5H2SO4(l) + 8KI(s) 🡪 H2S(g) + 4I2(g) + 4H2O(l) + 4K2SO4(s) 4. Cl- is the weakest reducing agent; it does not reduce H2SO4; Br- is a stronger reducing agent; it reduces S in H2SO4 from +6 to +4; I- is the strongest reducing agent; it reduces S in H2SO4 from +6 to -2 5. MnO2; it should also oxidise Br- and I- because these are stronger reducing agents than Cl- |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 3.2: Compare the reactions of concentrated sulphuric acid with the different halide ions** |
| 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 |

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| Can’t do this experiment? Watch this video instead: [www.youtube.com/watch?v=\_I5O5dYEdO4](http://www.youtube.com/watch?v=_I5O5dYEdO4) |

***Lesson 4 – What else do I need to know about halogens and halides?***

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| **Summary Activity 4.1: What can you remember about tests for halide ions?** |
| * Add HNO3 and then AgNO3; Cl- will give a white precipitate, Br- will give a cream precipitate and I- will give a yellow precipitate |

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| Image result for test icon**Test your knowledge 4.2: Group 7 – describing other reactions of halogens and halides** |
| 1. Mg + Cl2 🡪 MgCl2; Mg + Br2 🡪 MgBr2 2. Chlorine is more reactive than bromine as it attracts electrons more strongly 3. Cl2(g) + H2O(l) ==HCl(aq) + HClO(aq); HClO is an oxidising agent and can kill bacteria but is not harmful to humans in small concentrations 4. Cl2(g) + 2NaOH(aq) 🡪 NaCl(aq) + NaClO(aq) + H2O(l); NaClO is a strong oxidising agent and is used in domestic bleach 5. It changes from 0 to +1 and -1   Use blue litmus paper; chlorine turns it red and then bleaches it   1. Add HNO3 and then AgNO3; Cl- will give a white precipitate, Br- will give a cream precipitate and I- will give a yellow precipitate 2. 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?***

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| **Summary Activity 5.1: The special properties of water** |
| * simple molecular; water forms molecules (H2O) 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 NH4+, K+, Na+ or NO3-); simple molecules dissolve in water only if they are polar or can form hydrogen bonds; most giant covalent substances do not dissolve in water |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 5.2: Demonstrate the polarity of water** |
| 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 |

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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 5.3: Investigate the solubility of different substances in water** |
| 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 |

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| Image result for test icon**Test your knowledge 5.4: Understanding the Importance of Water** |
| 1. 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 2. 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 3. 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 4. Add a few drops to anhydrous copper sulphate; if the copper sulphate turns blue then water is present 5. 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?***

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| **Summary Activity 5.1: What do you know about hydrogen, oxygen and their compounds?** |
| * both simple molecular; both form diatomic molecules with very weak Van der Waal’s forces holding the molecules together * Eg Mg + 2HCl 🡪 MgCl2 + H2; 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 CaCO3 or another carbonate * OH- ions turn red litmus paper blue; they also give a pungent-smelling gas when warmed with ammonium chloride * CaO: CaO + 2HCl 🡪 CaCl2 + H2O; Na2O: Na2O + 2HCl 🡪 2NaCl + H2O * SO2 :SO2 + 2NaOH 🡪 Na2SO3 + H2O; Cl2O: Cl2O + 2NaOH 🡪 2NaClO + H2O * ZnO: ZnO + 2HCl 🡪 ZnCl2 + H2O; ZnO + 2NaOH 🡪 Na2ZnO2 + H2O |

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| https://image.freepik.com/free-icon/think-symbol-of-a-head-from-side-view-with-brain-shape-inside_318-61572.jpg **Activity 6.2: What is the best way to put out a fire?** |

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| Image result for test icon**Test your knowledge 6.3: Understanding the reactions of hydrogen, oxygen and their compounds** |
| 1. Insert a lit splint into the test tube – the hydrogen burns with a squeaky pop 2. H-O-O-H 3. 2H2O2 🡪 2H2O + O2; used in the laboratory preparation of oxygen 4. 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 5. C6H12O6 + 6O2 🡪 6CO2 + 6H2O; this is how all living things generate energy to function 6. Water 7. A fire extinguisher containing carbon dioxide or foam 8. Covalent oxides (ie oxides of non-metals) are acidic; ionic oxides (ie oxides of metals) are basic 9. Acidic: CO2, SO3, ClOH, HPO(OH)2, SiO2; basic: MgO, NaOH, FeO; amphoteric: Al2O3 10. CaO + SO2 🡪 CaSO3; 2NaOH + SiO2 🡪 Na2SiO3 + H2O; CO2 + MgO 🡪 MgCO3 |

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

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| **Summary Activity 7.1: What can you remember about sulphur dioxide?** |
| * SO2; it is acidic – SO2 + 2OH- 🡪 SO32- + H2O * It will turn blue litmus paper red and will turn acidified dichromate paper from orange to green * H2SO4; (i) H2SO4 + 2NaOH 🡪 Na2SO4 + 2H2O, acid-base or neutralisation; (ii) H2SO4 + 2NH3 🡪 (NH4)2SO4, acid-base or neutralisation; (iii) H2SO4 + 2NaCl 🡪 Na2SO4 + 2HCl, acid-base or acid-salt; (iv) 2H2SO4 + 2NaBr 🡪 Na2SO4 + Br2 + SO2 + 2H2O, redox; (v) 2H2SO4 + 2NaBr 🡪 Na2SO4 + Br2 + SO2 + 2H2O, redox; 5H2SO4(l) + 8NaI(s) 🡪 H2S(g) + 4I2(g) + 4H2O(l) + 4K2SO4(s), redox * Na2S, S2- * Add dilute HCl and then BaCl2 (aq), white precipitate seen |

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| Image result for test icon**Test your knowledge 7.2 – understanding sulphur and its compounds** |
| 1. The S can be oxidised from +4 to +6; SO2 is added to wine and sulphate (IV) ions are added to food to prevent oxidation 2. Low temperature gives higher yield of SO3 but slows down reaction so compromise temperature of 450 oC used; high pressure gives higher yield of SO3 and fast reaction but is expensive so 200 kPa used; V2O5 catalyst speeds up reaction 3. It is an acid and reacts with water in the air to form acid rain: SO3 + H2O 🡪 H2SO4 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 HNO3; (ii) it is used in the laboratory preparation of SO2; (iii) it is used to dry the SO2 produced in the laboratory preparation of SO2 5. Zn + S 🡪 ZnS (S reduced from O to -2) 6. ZnS + 2HCl 🡪 ZnCl2 + H2S (not a redox reaction) 7. 2ZnS + 3O2 🡪 2ZnO + 2SO2 (S oxidised from -2 to +4) this is the first stage in the extraction of Zn from its ore 8. See above 9. It smells of rotten eggs |

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

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| **Summary Activity 8.1: What can you remember about carbon, carbon dioxide and carbonate ions?** |
| * 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 * CO2 + H2OH2CO3H+ + HCO3-; CO2 + 2NaOH 🡪 Na2CO3 + H2O (CO2 is acidic) * Turns limewater milky and then colourless again * Bubbles on addition of HCl; CO2 given off; precipitate formed on addition of BaCl2 but precipitate dissolves in HCl |

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| Image result for test icon**Test your knowledge 8.2 – understanding carbon and its compounds** |
| 1. 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) 2. Coke is produced by the destructive distillation of coal; coal gas, coal tar and coal oil are also produced 3. Mixture of CO and H2, produced by heating coke at 700 oC in a limited supply of oxygen 4. CO2 + 2OH- 🡪 CO32- + H2O; CO32- + 2H+ 🡪 CO2 + H2O; 2HCO3- 🡪 CO32- + CO2 + H2O or HCO3- + H+ 🡪 CO2 + H2O and HCO3- + OH- 🡪 CO32- 5. Add dilute HCl to CaCO3; collect CO2 by downward delivery 6. It is not flammable and more dense than air 7. 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?**

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

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| Image result for test icon**Test your knowledge 9.1 – Understanding Nitrogen and Ammonia** |
| 1. 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; Ar boils first, then N2, so N2 collected at top of column 2. NH4Cl reacts with NaNO2; redox reaction; N2 not very soluble in water so can be collected over water 3. Ca(OH)2 reacts with NH4Cl; acid-base reaction; NH3 very soluble in water so not collected over water; less dense than air so collected by upward delivery 4. Low temperature gives higher yield of ammonia but slows down reaction so compromise temperature of 450 oC used; high pressure gives higher yield of ammonia and fast reaction so 25 MPa used; Fe catalyst speeds up reaction 5. To make fertilisers |

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

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| **Summary Activity 10.1: What can you remember about nitric acid and nitrates?** |
| * HNO3, NO3- * Nitric acid is a strong acid, so nitrate ions are neutral * HNO3 + NH3 🡪 NH4NO3 * Add almuminium powder, NaOH and heat; a pungent gas (NH3) is given off which turns damp red litmus paper blue |

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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 10.2: Decompose different nitrate salts** |
| 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  2NaNO3(s) 🡪 2NaNO2(s) + O2(g)  The N is reduced from + 5 to +3 and the O is oxidised from -2 to 0  2Ca(NO3)2(s) 🡪 2CaO(s) + 4NO2(g) + O2(g)  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 NO2 |
| Image result for test icon**Test your knowledge 10.3 – Nitric acid and nitrates** |
| 1. By adding concentrated sulphuric acid to potassium nitrate: H2SO4(l) + KNO3(s) 🡪 HNO3(g) + K2SO4(s); acid-base or acid-salt reaction 2. To make fertilisers (mainly ammonium nitrate) 3. 2NaNO3(s) 🡪 2NaNO2(s) + O2(g); 2Cu(NO3)2(s) 🡪 2CuO(s) + 4NO2(g) + O2(g) |

***Lesson 11 – How can we prepare and collect gases in the laboratory? (Part 1)***

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| **Summary Activity 11.1: Which reactions can be used to prepare different gases in the laboratory?** |
| * H2: Zn + 2HCl 🡪 ZnCl2 + H2 (redox or metal-acid); O2: 2H2O2 🡪 2H2O + O2 (redox or disproportionation); N2: NH4Cl + NaNO2 🡪 N2 + 2H2O + NaCl (redox); Cl2: MnO2 + 4HCl 🡪 MnCl2 + Cl2 + 2H2O (redox); HCl: H2SO4 + 2KCl 🡪 K2SO4 + 2HCl (acid-salt); SO2: either Na2SO3 + 2HCl 🡪 2NaCl + SO2 + H2O (acid-base) or Cu + 2H2SO4 🡪 CuSO4 + SO2 + 2H2O (redox); CO2: CaCO3 + 2HCl 🡪 CaCl2 + CO2 + H2O (acid-base); NH3: Ca(OH)2 + 2NH4Cl 🡪 CaCl2 + 2NH3 + 2H2O (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 |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 11.2: Prepare H2 gas in the laboratory** |
| **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 moldm-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   * Zn + 2HCl 🡪 ZnCl2 + H2 * 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: 2H2(g) + O2(g) 🡪 2H2O(l) |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 11.3: Prepare O2 gas in the laboratory** |
| **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   * 2H2O2(l) 🡪 2H2O + O2(g) * In this reaction the O is both oxidised (from -1 to O, 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 |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 11.4: Prepare N2 gas in the laboratory** |
| **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   * NH4Cl + NaNO2 🡪 N2 + 2H2O + NaCl * N in NH4+ is oxidised from -3 to 0 and N in NO2- is reduced from +3 to 0 * it is insoluble in water because it is non-polar |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 11.5: Prepare Cl2 gas in the laboratory** |
| **Equipment and chemicals needed:** a round-bottomed flask, two conical flasks, a dropping funnel, three bungs to fit the flasks, each with two holes, thee delivery tubes, a gas jar, Bunsen burner, tripod, gauze), 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   * MnO2 + 4HCl 🡪 MnCl2 + Cl2 + 2H2O * 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 * Cl2(g) + H2O(l) == HCl(aq) + HClO(aq); HCl contains H+ and turns litmus red; HClO is a bleaching agent and turns it white |

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| Image result for test icon**Test your knowledge 11.6: Describing the collection of different gases (Part 1)** |
| 1. (i) H2: Zn + 2HCl 🡪 ZnCl2 + H2 (redox or metal-acid); (ii) O2: 2H2O2 🡪 2H2O + O2 (redox or disproportionation); (iii) N2: NH4Cl + NaNO2 🡪 N2 + 2H2O + NaCl (redox); (iv) Cl2: MnO2 + 4HCl 🡪 MnCl2 + Cl2 + 2H2O (redox) 2. H2, O2 and N2 have limited solubility in water so are collected over water; Cl2 reacts with water and is denser than air so is collected by downward delivery 3. H2 burns with a squeaky pop; O2 relights a glowing splint; Cl2 turns blue litmus red and then white |

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

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 12.1: Prepare HCl gas in the laboratory** |
| **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 H2SO4 , filter paper soaked in conc NH3 (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 * H2SO4(l) + 2KCl(s) 🡪 K2SO4(s) + 2HCl(g) * HCl(g) 🡪 H+(aq) + Cl-(aq) (H+ turns the blue litmus red) * NH3(g) + HCl(g) 🡪 NH4Cl(s) (NH4Cl is the white smoke) |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 12.2: Prepare SO2 gas in the laboratory (Part 1)** |
| **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 moldm-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   * Na2SO3(s) + 2HCl(aq) 🡪 2NaCl(aq) + SO2(g) + H2O(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 * SO2(g) + H2O(l) == H2SO3(aq) ==2H+(aq) + SO32-(aq); H+ and turns litmus red * SO2 is a reducing agent; it gets oxidised to SO42- ions and reduces dichromate ions to Cr3+ |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 12.3: Prepare SO2 gas in the laboratory (Part 2)** |
| **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   * Cu + 2H2SO4 🡪 CuSO4 + SO2 + 2H2O * S reduced from +6 to +4; Cu oxidised from 0 to +2 |
| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 12.4: Prepare CO2 gas in the laboratory** |
| **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   * CaCO3(s) + 2HCl(aq) 🡪 CaCl2(aq) + CO2(g) + H2O(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; CO2(g) + Ca(OH)2(aq) 🡪 CaCO3(s) + H2O(l); CaCO(s) + CO2(g) + H2O(l) 🡪 Ca(HCO3)2(aq) * Carbon dioxide extinguishes a lit splint as it is not flammable |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 12.5: Prepare NH3 gas in the laboratory** |
| **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   * Ca(OH)2(s) + 2NH4Cl(s) 🡪 CaCl2(s) + 2NH3(g) + 2H2O(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 |

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| Image result for test icon**Test your knowledge 12.6: Describing the collection of different gases (Part 2)** |
| 1. (i) HCl: H2SO4 + 2KCl 🡪 K2SO4 + 2HCl (acid-salt); (ii) SO2: either Na2SO3 + 2HCl 🡪 2NaCl + SO2 + H2O (acid-base) or Cu + 2H2SO4 🡪 CuSO4 + SO2 + 2H2O (redox); (iii) CO2: CaCO3 + 2HCl 🡪 CaCl2 + CO2 + H2O (acid-base); (iv) NH3: Ca(OH)2 + 2NH4Cl 🡪 CaCl2 + 2NH3 + 2H2O (acid-base) 2. HCl, SO2, CO2 and NH3 are highly soluble in water so cannot be collected over water; HCl, SO2 andCO2 are denser than air so collected by downward delivery; NH3 is less dense than air so collected by upward delivery 3. 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 |

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

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| Image result for test icon**Test your knowledge 13.1 – understanding rocks** |
| 1. Mineral – substance of fixed composition found in the earth’s crust; rock – mixture of minerals 2. 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 3. Breaking down of rocks into smaller particles 4. 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 |

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

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| Image result for test icon  **14.1 END-OF-UNIT QUIZ**  **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) Cl2 + 2NaOH 🡪 NaCl + NaClO + H2O   (b) Cl2 + 2NaBr 🡪 2NaCl + Br2   1. 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 2. N2 + 3H22NH3; 450 oC (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 3. (a) 2Ca(NO3)2 🡪 2CaO + 4NO2 + O2   (b) 2NaNO3 + 2NaNO2 + O2  (c) 2CuS + 3O2 🡪 2CuO + 2SO2   1. Peat compressed into lignite and then into coal; over millions of years it is further compressed at high temperatures into anthracite 2. 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 3. Breaking down of rocks into smaller pieces (hydration, hydrolysis, carbonation and oxidation) |