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| **UNIT 6**  **REDOX REACTIONS**  **Answers** |

***Lesson 1 – What is oxidation, what is reduction and what are oxidation numbers?***

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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**Thinkabout Activity 1.1: What is oxidation and what is reduction?** |
| * 2Mg + O2 🡪 2MgO * Mg loses its two valence electrons * Mg is oxidised because it “gains oxygen”; it is also oxidised because it “loses electrons” * Mg 🡪 Mg2+ + 2e; O2 + 4e 🡪 2O2- |

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| **Image result for test iconTest your knowledge 1.2: Deducing Oxidation Numbers** | | |
| 1. +4 2. -2 3. +4 4. +6 5. +5 6. +3 7. 0 | 1. +2 2. +1 3. +5 4. +1 5. -1 6. -1 7. +3- | 1. +5 2. +7 3. +2 4. +8/3 5. +2.5 6. +2 |

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| **Image result for test iconTest your knowledge 1.3: Using oxidation numbers to identify oxidation and reduction** |
| 1. +4 to +2, so reduction 2. -1 to 0, so oxidation 3. +1 to -1, so reduction 4. +2 to +2.5, so oxidation 5. +5 to +2, so reduction |

***Lesson 2 – How can I use oxidation numbers to name inorganic compounds?***

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| **Image result for test iconTest your knowledge 2.1: Naming binary inorganic compounds** | | | |
| 1. copper (II) oxide 2. copper (I) oxide 3. lead (IV) oxide 4. calcium (II) nitride 5. carbon (II) oxide 6. nitrogen (IV) oxide 7. nitrogen (III) chloride | 1. lead (II) sulphide 2. hydrogen (I) nitride 3. oxygen (II) fluoride 4. hydrogen (I) telluride 5. uranium (VI) fluoride 6. aluminium (III) hydride 7. iron (II) oxide | | 1. iron (III) oxide 2. manganese (IV) oxide 3. carbon (IV) chloride 4. chlorine (VII) oxide 5. tin (IV) chloride 6. tin (II) chloride 7. sulphur (IV) oxide |
| **Image result for test iconTest your knowledge 2.2: Naming non-binary inorganic compounds**  Give the IUPAC name of the following compounds: | | | |
| 1. sodium sulphate (IV) 2. potassium chlorate (V) 3. sodium chlorate (I) 4. copper (II) carbonate 5. magnesium nitrate | | 1. sodium nitrate (III) 2. ammonium sulphate 3. potassium manganate (VIII) 4. potassium chromate (VI) 5. potassium hexafluoriplatinate (IV) | |
| Note: the oxidation number not been used with the cation unless it is a d-block metal and has a number of stable oxidation numbers  Note: the oxidation number has been omitted with very common anions (eg sulphate, nitrate, carbonate) but is always used with lesson common anions or if the oxidation number is not the most common one (eg sulphate (IV) and nitrate (III) | | | |

***Lesson 3 – What are half-equations and how can we construct them?***

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| **Image result for test iconTest your knowledge 3.1: Writing half-equations** | |
| 1. PbO2 + 2H+ + 2e- 🡪 Pb2+ + 2H2O 2. 2Cl- 🡪 Cl2 + 2e- 3. 2S2O32- 🡪 S4O62- + 2e- 4. 2IO3- + 12H+ + 10e- 🡪 I2 + 6H2O 5. 2I- 🡪 I2 + 2e- | 1. ClO- + 2H2O 🡪 ClO3- + 4H+ + 4e- 2. ClO- + 2H+ + 2e- 🡪 Cl- + H2O 3. H2SO4 + 2H+ + 2e- 🡪 SO2 + 2H2O 4. 2Br- 🡪 Br2 + 2e- 5. H2SO4 + 6H+ + 6e- 🡪 S + 4H2O |

***Lesson 4 – What are redox reactions, and what are oxidising and reducing agents?***

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| **Image result for test iconTest your knowledge 4.1: Writing equations for redox reactions** | |
| 1. PbO2 + 4H+ + 2Cl- 🡪 Pb2+ + Cl2 + 2H2O 2. 4Al3+ + 6O2- 🡪 3O2 + 4Al 3. 2S2O32- + I2 🡪 S4O62- + 2I- 4. IO3- + 5I- + 6H+ 🡪 3I2 + 3H2O | 1. 3ClO- 🡪 ClO3- + 2Cl- 2. H2SO4 + 2Br- +2H+ 🡪 SO2 + Br2 + 2H2O 3. H2SO4 + 6H+ + 6I- 🡪 3I2 + S + 4H2O 4. ClO- + 2H+ + 2I- 🡪 Cl- + I2 + H2O |

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| **Image result for test iconTest your knowledge 4.2: Identifying oxidising and reducing agents in redox reactions** | |
| 1. oxidising agent is PbO2; reducing agent is Cl- 2. oxidising agent is Al3+; reducing agent is O2- 3. oxidising agent is I2; reducing agent is S2O32- 4. oxidising agent is IO3-; reducing agent is I- | 1. oxidising agent is ClO-; reducing agent is ClO- 2. oxidising agent is H2SO4; reducing agent is Br- 3. oxidising agent is H2SO4; reducing agent is I- 4. oxidising agent is ClO-; reducing agent is I- |

***Lesson 5 – What are the some common examples of redox reactions?***

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| **Practical 5.1: Prepare a sample of zinc sulphate from zinc and sulphuric acid** |
| Equipment needed per group: 100 cm3 beaker, tripod, gauze, Bunsen burner, thermometer, 20 cm3 of 0.5 moldm-3 H2SO4, 1 g zinc granules, 2 pieces filter paper, funnel, spatula, 100 cm3 conical flask, label, access to of 0.5 moldm-3 H2SO4 (20 cm3 per group), access to zinc granules (1 g per group), weighing boat, mass balance  Make sure students do not overheat the mixture   * Moles of H2SO4 = 20/1000 x 0.5 = 0.01 * Moles of Zn = 1/65.4 = 0.015 * Excess zinc can be easily removed by filtration; excess H2SO4 cannot easily be removed |

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| **Image result for test iconTest your knowledge 5.2: Analysing common redox reactions** |
| 1. Zn + 2HCl 🡪 ZnCl2 + H2 or Zn + 2H+ 🡪 Zn2+ + H2 (o = Zn; r = H) 2. 2Al + 6HNO3 🡪 2Al(NO3)3 + 3H2 or 2Al + 6H+ 🡪 2Al3+ + 3H2 (o = Al; r = H) 3. ZnSO4 + Mg 🡪 MgSO4 + Zn or Zn2+ + Mg 🡪 Mg2+ + Zn (o = Mg; r = Zn) 4. 2AgNO3 + Cu 🡪 Cu(NO3)2 + 2Ag or 2Ag+ + Cu 🡪 Cu2+ + 2Ag (o = Cu; r = Ag) 5. 4Ag + O2 🡪 2Ag2O (o = Ag; r = O) 6. Fe2O3 + 3CO 🡪 2Fe + 3CO2 (o = C; r = Fe) 7. TiCl4 + 2Mg 🡪 Ti + 2MgCl2 (o = Mg; r = Ti) |

***Lesson 6 – How can we identify oxidising and reducing agents?***

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| **Summary Activity 6.1: What can you remember about qualitative analysis?** |
| * Identification of ions or molecules by simple tests * CO32-: add HCl (aq) and observe fizzing, or add CaCl2 (aq) and observe white precipitate, then add HCl (aq) and observe fizzing as the precipitate dissolves; gas should be odourless * SO32-: add HCl (aq) and observe fizzing, or add CaCl2 (aq) and observe white precipitate, then add HCl (aq) and observe fizzing as the precipitate dissolves; gas should smell like burning matches * NH4+: add NaOH (aq) and warm; pungent gas should be given off * H+: add CaCO3(s) and observe fizzing, or add blue litmus paper – it will turn red * OH-: add NH4Cl (aq) and warm; pungent gas should be given off or add red litmus paper – it will turn blue * CO2: turns limewater milky and then colourless; no smell * SO2: turns limewater milky and then colourless; smell of burning matches; will turn damp blue litmus paper red * NH3: pungent smell, will turn damp red litmus paper blue, will give white smoke with filter paper soaked in concentrated HCl * HCl: will turn damp blue litmus paper red, will give white smoke with filter paper soaked in concentrated NH3 |

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| **Practical 6.2: Test for oxidising and reducing agents in solution** |
| Chemicals required per group: access to bottles of 0.05 moldm-3 Fe2(SO4)3 (labelled A), 0.1 moldm-3 FeSO4 (labelled B), 0.1 moldm-3 KNO3 (labelled C), 0.1 moldm-3 Na2SO3 (labelled D), 1 moldm-3 NaOH, 1 moldm-3 HCl, 0.02 moldm-3 K2Cr2O7 in 1 moldm-3 H2SO4, 0.02 moldm-3 KMnO4 in 1 moldm-3 H2SO4, each with its own dropping pipette (approx 3 cm3 of each needed per group), access to pots of Na2SO3 and Al powder, each with a spatula (approx 1 g per group), access to red litmus paper, access to dichromate paper  Apparatus required per group: 12 test tubes, 1 test tube rack, Bunsen burner, tongs   * Solution A should give a brown colour with KI * Solutions B and D should decolorise KMnO4 and turn K2Cr2O7 green * Solution C should not give a positive test for the above but will give off a pungent gas when warmed with NaOH and Al; the gas should turn red litmus blue * Na2SO3 should fizz on addition of HCl and the gas evolved should turn dichromate paper green |

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| **Image result for test iconTest your knowledge 6.3: Describing qualitative analysis using redox reactions** |
| 1. Warm with NaOH and Al powder; pungent-smelling gas should be evolved which turns red litmus blue 2. Add acidified KMnO4; SO32- will decolorise it but CO32- will not OR add acidified K2Cr2O7; SO32- will turn it green it but CO32- will not 3. Add HCl; SO32- sample will evolve gas with burning-match smell which turns dichromate paper green; CO32- sample will evolve odourless gas which has no effect on dichromate paper 4. SO2 has burning-match smell and turns dichromate paper green; CO2 is odourless and has no effect on dichromate paper 5. Fe2+ decolorises acidified KMnO4 but does not react with KI; Fe3+ turns KI brown but has no effect on acidified KMnO4 |

***Lesson 7 – What is a Galvanic cell?***

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| **Practical 7.1: Build a simple electrochemical cell** |
| Equipment needed per group: 2 x 100 cm3 beakers, 1 copper strip, 1 zinc strip, 1 strip filter paper (1 x 15 cm), 2 crocodile clips, 2 electrical wires, 1 voltmeter, access to 1 moldm-3 CuSO4, 1 moldm-3 ZnSO4 (50 cm3 per group), saturated KNO3 (10 cm3 per group), disposable gloves  The voltmeter should read around 1.1 V; either +1.1 V (if the Cu electrode is on the RHS) or -1.1 V (if the Cu electrode is on the left hand side; the Cu is the positive electrode  Cu2+ + 2e 🡪 Cu; Zn 🡪 Zn2+ + 2e |

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| **Image result for test iconTest your knowledge 7.2: Understanding how Galvanic cells work** |
| 1. Voltmeter reading -ve so positive electrode is on LHS, and is therefore the zinc electrode, so magnesium electrode is negative 2. Zn electrode +ve so reduction: Zn2+ + 2e 🡪 Zn; Mg electrode -ve so oxidation: Mg 🡪 Mg2+ + 2e 3. Zn2+ + Mg 🡪 Zn + Mg2+ 4. Electrons move from Mg (oxidised) to Zn (reduced) 5. Sulphate ions move from Zn2+ (which is decreasing in concentration) to Mg2+ (which is increasing in concentration) |

***Lesson 8 – What are electrode potentials and how can we measure them?***

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| **Image result for test iconTest your knowledge 8.1: Using the Conventional Representation of Cells** |
| 1. Mg|Mg2+||Zn2+|Zn 2. Zn|Zn2+||Fe2+|Fe 3. Zn|Zn2+||Ag+|Ag 4. Al|Al3+||Pb2+|Pb 5. Mg|Mg2+||Al3+|Al |

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| **Image result for test iconTest your knowledge 8.2: Using standard electrode potentials** |
| 1. (i) Cu|Cu2+||Ag+|Ag; emf = +0.46 V; 2Ag+ + Cu 🡪 Cu2+ + 2Ag   (ii) Zn|Zn2+||Pb2+|Pb; emf = +0.63 V; Pb2+ + Zn 🡪 Zn2+ + Pb   1. Al|Al3+||Fe2+|Fe; emf = +1.22 V; 3Fe2+ + 2Al 🡪 2Al3+ + 3Fe 2. Mg|Mg2+||Al3+|Al; emf = +0.71 V; 2Al3+ + 2Mg 🡪 2Mg2+ + 2Al 3. Pb|Pb2+||Ag+|Ag; emf = +0.93 V; 2Ag+ + Pb 🡪 Pb2+ + 2Ag 4. (i) emf -ve so oxidation on RHS so V 🡪 V2+ + 2e, so Cu2+ + 2e 🡪 Cu: Cu2+ + V 🡪 V2+ + Cu   (ii) emf = Er - El so -1.46 = V2+|V – 0.34 so V2+|V = -1.46 + 0.34 = -1.12 V |

***Lesson 9 – What are the different types of Galvanic cell?***

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| **Image result for test iconTest your knowledge 9.1: Understanding Applications of Galvanic Cells** |
| 1. A primary cell is non-rechargeable (eg most typical batteries); a secondary cell is rechargeable (eg a lead-acid battery) 2. It is rechargeable and can withstand a large current for a short time 3. The electrolyte is in the form of a paste rather than a liquid or solution; they can be turned around and used in any orientation 4. Reaction between fuel and oxygen; fuel is constantly entering the cell at one electrode and oxygen at the other; the products leave the electrode as they are made; they are more effiecient than combustion engines |

***Lesson 10 – What is electrolysis?***

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| **Image result for test iconTest your knowledge 10.1: Describing electrolysis** |
| 1. (i) cathode: Na+ + e- 🡪 Na; anode: 2Cl- 🡪 Cl2 + 2e-; overall: 2NaCl 🡪 2Na + Cl2   (ii) cathode: Al3+ + 3e- 🡪 Al; anode: 2O2- 🡪 O2 + 4e-; overall: 2Al2O3 🡪 4Al + 3O2  (iii) cathode: 2H2O + 2e- 🡪 H2 + 2OH-; anode: 2Cl- 🡪 Cl2 + 2e-; overall: 2H2O + 2Cl- 🡪 H2 + Cl2 + 2OH-  (iv) cathode: 2H+ + 2e- 🡪 H2; anode: 2H2O 🡪 O2 + 4e- + 4H+; overall: 2H2O 🡪 2H2 + O2  (v) cathode: Cu2+ + 2e- 🡪 Cu; anode: 2H2O 🡪 O2 + 4e- + 4H+; overall: 2Cu2+ + 2H2O 🡪 2Cu + O2 + 4H+  (b) cathode: Cu2+ + 2e- 🡪 Cu; anode: Cu 🡪 Cu2+ + 2e-; used in purification of copper |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 10.2: Electrolyse aqueous solutions** |
| **Equipment needed:** one power pack able to produce a variable power supply up to 12 V, two electrical wires, two crocodile clips, two graphite electrodes, preferably with a support to keep them parallel, three 100 cm3 beakers, access to 1 moldm-3 solutions of NaOH, CuCl2 and KI   * Electrolysis of NaOH: bubbles at both electrodes, more concentrated at cathode; gases could be collected and confirmed as H2 at cathode (burns with a squeaky pop) and O2 at anode (relights glowing splint); 2H2O 🡪 2H2 + O2 * Electrolysis of CuCl2: bubbles at anode, layer of copper at cathode; gas at anode should bleach litmus paper; CuCl2 🡪 Cu + Cl2 * Electrolysis of KI: bubbles at cathode, brown solution of I2 at anode; gas at cathode could be collected and confirmed as H2 (burns with a squeaky pop); solution should turn red litmus blue as NaOH is being formed; 2KI + 2H2O 🡪 H2 + I2 + 2KOH   Note: the tests for hydrogen, oxygen and chlorine gases are covered in Unit 12 – Non-metals and their compounds |

***Lesson 11 – How can we predict the quantity of each substance produced during electrolysis?***

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| **Summary Activity 11.1: Units of charge and amount of substance** |
| * Pipette: very accurate but can only deliver one volume * Volumetric flask: very accurate but can only store one volume * Burette: slightly less accurate than a pipette but can deliver any volume up to 50 cm3 * Measuring cylinder: not accurate * Pipettes and burettes are most useful for carrying out titrations |

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| **Image result for test iconTest your knowledge 11.2: Calculating the quantities produced during electrolysis** |
| 1. Electrolyse dilute NaCl or dilute H2SO4 using electrodes fully submerged; place inverted measuring cylinders filled with water above the electrodes and collect the gases produced at each electrode; the gas at the cathode will have twice the volume of the gas at the anode 2. 20/58.5 = 0.34 moles of NaCl; each NaCl needs 1 electron for separation so 0.34 F needed = 33,000 C 3. 2000/96500 = 0.02 moles; which will produce 0.01 moles of H2 and 0.05 moles of O2; V = nRT/P = 2.6 x 10-4 m3 = 0.26 dm3 of hydrogen and 0.13 dm3 of oxygen 4. 1000 g of Al2O3 = 0.98 moles; each Al2O3 requires 6e for electrolysis so 0.98 x 6 = 5.88 F = 568,000 C 5. 1000/96500 = 0.0104 F; each Cu requires 2e so moles of Cu = 0.00518 so mass = 0.00518 x 63.5 = 0.33 g 6. n = PV/RT = (100000 x 100 x 10-6)/(8.31 x 298) = 0.00403; 2e required per H2 so 0.00808 F required; 0.00808 x 96500 = 779 C 7. 2/107.9 = 0.0185 moles Ag; each Ag needs 1 e so 0.0185 C required = 1790 C |

***Lesson 12 – What is rusting and how can we prevent it?***

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| **Practical 12.1: Demonstrate that oxygen and water are both required for rusting** |
| Equipment needed per group: 3 test tubes, 3 bungs which fit the test tubes, 1 test tube rack, three small nails (2 - 4 cm), access to anhydrous CaCl2 and a spatula (1 g per group), access to vegetable oil with dropping pipette (1 cm3 per group), 5 cm3 of recently boiled water, access to distilled water   * The nail in tubes B and C should show no signs of rusting after one week, but the nail in tube A should show clear signs of rusting |

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| **Practical 12.2: Demonstrate that acids, bases, salts and heating increase the rate of rusting** |
| Equipment needed per group: 5 test tubes, one test tube rack, 5 small nails (2 – 4 cm), access to distilled water, access to 0.1 moldm-3 solutions of HCl and NaOH, each with dropping pipette (2 cm3 per group) and NaCl with dropping pipette (4 cm3 per group), access to a fridge   * The test tubes containing HCl, NaOH and NaCl should all have rusted significantly more than the test tube containing water only; the acid test tube may not appear rusty as rust dissolves in acid; the solution may appear green/orange; the test tube containing NaCl in the fridge should rust less than the test tube containing NaCl in a warm place |

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| **Image result for test iconTest your knowledge 12.3: Understanding Rusting** |
| 1. 4Fe + 6H2O + 3O2 🡪 4Fe(OH)3; Fe oxidised from 0 to 3; O reduced from 0 to -2 2. Oxygen (air) and water 3. Salt, acids, alkalis (ie electrolytes), heat 4. Oiling, painting, galvanisation 5. Attaching pieces of a more reactive metal to the iron surface; the more reactive metal is oxidised preferentially; zinc can provide sacrificial protection but tin cannot |

***Lesson 13 – What have I learned and understood about redox reactions?***

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| **13.1 END-OF-TOPIC QUIZ**  **UNIT 6 – ACIDS, BASES AND SALTS**  Image result for test icon |
| 1. (a) +1, sodium chlorate (I); (b) +4, potassium sulphate (IV); (c) +7, potassium manganate (VII) 2. (a) reduction: ClO- + 2H+ + 2e- 🡪 Cl- + H2O; (b) oxidation: SO32- + H2O 🡪 SO42- + 2H+ + 2e-; (c) reduction: MnO4- + 8H+ + 5e- 🡪 Mn2+ + 4H2O; (d) oxidation: Fe2+ 🡪 Fe3+ + e-; (e) reduction: H2SO4 + 8H+ + 8e- 🡪 H2S + 4H2O; (f) oxidation: 2I- 🡪 I2 + 2e- 3. (a) H2SO4 + 8H+ + 8I- 🡪 H2S + 4I2 + 4H2O; (b) MnO4- + 8H+ + 5Fe2+ 🡪 Mn2+ +4H2O + 5Fe3+; (c) ClO- +SO32- 🡪 Cl- + SO42- 4. Add excess Mg to dilute HNO3; filter off excess magnesium; heat solution until 75% evaporated; then leave until crystals form; dry crystals in filter paper; Mg + 2H+ 🡪 Mg2+ + H2; Mg oxidised from O to +2; H reduced from +1 to 0 5. (a) Warm with dilute NaOH and Al powder; a pungent gas will be given off which turns red litmus blue and/or gives a white smoke when in contact with HCl; (b) Add dilute HCl; a gas will be given off which smells like burning matches and which turns dichromate paper from orange to green 6. (a) Zn|Zn2+||Ni2+|Ni; (b) +0.62 V, Zn + Ni2+ 🡪 Zn2+ + Ni 7. (a) cathode: 2H2O + 2e- 🡪 H2 + 2OH-; anode: 2Cl- 🡪 Cl2 + 2e-; (b) 2H2O + 2Cl- 🡪 H2 + Cl2 + 2OH-; 2000/96500 = 0.0207 F; 1 H2 needs 2e so 0.0104 mol H2 produced; V = nRT/P = 0.0104 x 8.31 x 298 / 100000 = 2.57 x 10-4 m3 = 0.257 dm3 8. Make it the cathode of an electrolytic cell containing AgNO3 as the electrolyte and with an anode made of silver, and pass a current through 9. (a) 4Fe + O2 + 2H2O 🡪 4Fe(OH)3; (b) dissolved electrolytes and heat; (c) attaching the iron to a piece of a more reactive metal; (d) oiling, greasing, painting |