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| UNIT 9**METALS AND THEIR COMPOUNDS****Answers** |

***Lesson 1 – What are metals?***

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| **Summary Activity 1.1: What can you remember about metals?** |
| * A metal is a substance (usually an element) which contains metallic bonding
* Metallic bonding is the attraction between a lattice of cations and a sea of delocalised electrons
* Metallic bonding is quite strong so metals often have high melting points; the delocalised electrons make them good conductors of electricity; the metal ions can move past each other without disrupting the metallic bonding, so metals tend to be malleable and ductile
* A non-metal is an element which contains covalent bonding
* Non-metals either have simple molecular structures - small groups of atoms held together by covalent bonds (called molecules) and weak Van der Waal’s forces between the molecules, or giant covalent structures (lattice of atoms held together by covalent bonds
* Electropositive atoms do not hold on to their electrons strongly and allow their valence electrons to be delocalised; electronegative atoms hold on to their electrons strongly and form covalent bonds instead
* An alloy is a mixture of atoms held together by metallic bonds; the major component of the mixture must be a metal (eg brass, bronze, steel, solder)
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| **Image result for test iconTest your knowledge 1.2: Classifying metals, non-metals and metalloids** |
| 1. Eg sodium, calcium, magnesium, potassium
2. Eg aluminium, tin, lead
3. Eg copper, iron, zinc
4. Eg boron, silicon
5. Eg oxygen, bromine, neon
6. Electronegativity increases across a Period, so the attraction to bonding electrons increases, so atoms become less likely to allow bonding electrons to delocalise
7. Electronegativity decreases down a Geriod, so the attraction to bonding electrons decreases, so atoms become more likely to allow bonding electrons to delocalise
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***Lesson 2 – What are the physical properties of metals?***

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| **Image result for test iconTest your knowledge 2.1: Describing physical properties of metals** |
| 1. Delocalised electrons are free to move
2. cations can move around without breaking metallic bonds
3. Mg2+ is smaller than Na+ and is more highly charged, so it attracts delocalised electrons more strongly, so the metallic bonds are stronger and more energy is needed to break them
4. K+ is larger than Na+, so it attracts delocalised electrons less strongly, so the metallic bonds are weaker and less energy is needed to break them
5. Al3+ is smaller than Mg2+ and is more highly charged, so it attracts delocalised electrons more strongly, so the metallic bonding is stronger
6. Iron has a larger atomic mass than aluminium
7. Iron has more unpaired electrons than copper
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***Lesson 3 – How do s and p-block metals react with air, water and acids?***

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| **Summary Activity 3.1: What can you remember about redox reactions of metals?** |
| * Eg 2Mg + O2 🡪 2MgO or 4Al + 3O2 🡪 2Al2O3 (metal oxidised, O reduced)
* Eg Mg + 2H+ 🡪 Mg2+ + H2 or Zn + 2H+ 🡪 Zn2+ + H2 (metal oxidised, H+ reduced)
* Eg Zn + Cu2+ 🡪 Zn2+ + Cu (Zn oxidised, Cu2+ reduced)
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| **Image result for test iconTest your knowledge 3.2: Describing chemical properties of s and p-block metals** |
| 1. **2K + 2H2O 🡪 2KOH + H2; Ca + 2H2O 🡪 Ca(OH)2 + H2; fizzing, the metal dissolves, reaction is faster with K; redox reaction, K is larger than Ca and has fewer protons, so the attraction between the nucleus and outer electrons is weaker and it loses its electrons more easily**
2. **2Na + O2 🡪 Na2O2; 2Ca + O2 🡪 2CaO; sodium forms a peroxide, calcium forms an oxide**
3. **To prevent them from reacting with air or water**
4. **Mg + 2HCl 🡪 MgCl2 + H2 or Mg + 2H+ 🡪 Mg2+ + H2; 2Al + 6HCl 🡪 AlCl3 + 3H2 or 2Al + 6H+ 🡪 2Al3+ + 3H2; redox reaction; Mg is larger than Al and has fewer protons, so the attraction between the nucleus and outer electrons is weaker and it loses its electrons more easily**
5. **Al forms a very stable oxide layer on its surface which protects it from further reaction**
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***Lesson 4 - How do d-block metals react with air, water and acids?***

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| **Summary Activity 4.1: What can you remember about d-block metals?** |
| * Iron rusts when is is oxidised by O2 and H2O to Fe(OH)3; the rust does not stick to the surface to the iron but flakes off, exposing the iron underneath to further reaction
* By painting, greasing, galvanising, sacrificial protection with a more reactive metal
* By electrolysis of CuSO4 using copper electrodes; the copper on the impure anode dissolves (Cu 🡪 Cu2+ + 2e and pure copper is deposited at the cathode: Cu2+ + 2e 🡪 Cu
* Cu: 1s22s22p63s23p64s13d10; Fe: 1s22s22p63s23p64s23d6
* Cu2+: 1s22s22p63s23p63d9; Fe2+: 1s22s22p63s23p63d6; Fe3+: 1s22s22p63s23p63d5
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| **Image result for test iconTest your knowledge 4.2: Describing chemical properties of d-block metals** |
| 1. **because they do not always lose all of their valence d-electrons; the number of electrons they lose can vary**
2. **d-block metal: has valence s electrons and d electrons but no p-electrons; transition metal: forms at least one stable ion with a partially filled d-orbital (not all d-block metals are transition metals)**
3. **they can change their oxidation state so can accept and donate electrons; eg Fe in Haber process**
4. 4Fe + 6H2O + 3O2 🡪 4Fe(OH)3
5. **Copper does not react with water; gold does not react with air or water so stays shiny for a long time**
6. **Fe + H2SO4 🡪 FeSO4 + H2**
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***Lesson 5 – How can we compare the reactivity of different metals?***

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| **Summary Activity 5.1: Which types of redox reaction involve metals?** |
| * 2Na + 2H2O 🡪 2NaOH + H2; Na oxidised from 0 to +1; H reduced from +1 to 0; zinc less reactive than sodium and cannot displace hydrogen from water
* Zn + 2HCl 🡪 ZnCl + H2; Zn oxidised from 0 to +2; H reduced from +1 to 0; copper less reactive than zinc and cannot displace hydrogen from acids
* Zn + CuSO4 🡪 Cu + ZnSO4; Zn oxidised from 0 to +2; Cu reduced from +2 to 0; copper less reactive than zinc so cannot displace zinc from its compounds
* ZnO + C 🡪 Zn + CO; C oxidised from 0 to +2; Zn reduced from +2 to 0; aluminium more reactive than carbon so cannot be displaced from its compounds by carbon
* CuO + H2 🡪 Cu + H2O; H oxidised from 0 to +1; Cu reduced from +2 to 0; zinc more reactive than hydrogen so cannot be displaced from its compounds by hydrogen
* These are all examples of metal displacement reactions
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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 5.2: Compare the reactivity of different metals** |
| Chemicals needed per group: 12 test tubes, three 10 cm3 measurin cylinders, one test tube rack, one thermometer access to 0.5 moldm-3 solutions of CuSO4, ZnSO4, MgSO4 and FeSO4 (5 cm3 per group), each bottle with its own dropping pipette; access to powdered samples of Zn, Fe, Cu and Mg, each with its own spatulaSigns of reaction will include: a temperature rise which can be large, bubbles, a change in colour of the solution or of the powder; the largest temperature change will be with Mg and CuSO4; the reactions of Zn with CuSO4 and Mg with FeSO4 may also be vigorous.CuSO4 with Mg, Zn and Fe: CuSO4 + Mg 🡪 MgSO4 + Cu; CuSO4 + Zn 🡪 ZnSO4 + Cu; CuSO4 + Fe 🡪 FeSO4 + CuFeSO4 with Mg and Zn: FeSO4 + Mg 🡪 MgSO4 + Fe; FeSO4 + Zn 🡪 ZnSO4 + FeZnSO4 with Mg: ZnSO4 + Mg 🡪 MgSO4 + ZnMg most reactive as it displaces Cu, Fe and Zn from their salts; then Zn which displaces Fe and Cu but not Mg from their salts; then Fe which can only displace Cu from their salts, then Cu which cannot displace any of the other metals from their salts |

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| **Image result for test iconTest your knowledge 5.3: Understanding metal displacement reactions** |
| 1. **no reaction; (b) Mg + CuSO4 🡪 MgSO4 + Cu; (c) no reaction; (d) no reaction; (e) Zn + CuSO4 🡪 ZnSO4 + Cu; (f) no reaction; (g) Fe2O3 + 3C 🡪 Fe2O3 + 3CO; (h) SnO2 + 2C 🡪 Sn + 2CO**
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***Lesson 6 - How are metals extracted from their ores?***

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| **Summary Activity 6.1: Electrolytic Processes** |
| * Cathode: Al3+ + 3e- 🡪 Al; anode 2O2- 🡪 O2 + 4e-
* Cathode: Cu2+ + 2e 🡪 Cu; anode: Cu 🡪 Cu2+ + 2e-
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| **Image result for test iconTest your knowledge 6.2: Understanding the Extraction of Metals** |
| 1. Reactivity, required purity, cost of process
2. Lots of energy needed to melt the cryolite and for the electricity
3. The melting point of cryolite is lower than the melting point of pure aluminium oxide
4. Al3+ + 3e 🡪 Al (at the cathode) and 2O2- 🡪 O2 + 4e (at the anode)
5. Anodes react with oxygen C + O2 🡪 CO2
6. C + O2 🡪 CO2; C + CO2 🡪 2CO
7. SnO2 + 2CO 🡪 Sn + 2CO2
8. Fe2O3 + 3CO 🡪 2Fe + 3CO2
9. It helps remove the main impurity SiO2; CaCO3 decomposes to produce CaO, which reacts with SiO2 to produce CaSiO3, which can be removed
10. Oxygen is bubbled through the molten iron; the oxygen removes the C as CO2: C + O2 🡪 CO2
11. Al is more reactive than C and cannot be reduced from its oxide by C or CO
12. CaSiO3 is used in road-building
13. Fe is magnetic so can be separated from other scrap using a magnet
14. It is present in low concentrations and is difficult to obtain in pure form
15. It reacts with oxygen in the presence of cyanide ions to form a soluble compound; this compound is converted back to gold by reaction with carbon
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***Lesson 7 – Why are metals and their compounds useful (part I)?***

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| **Image result for test iconTest your knowledge 7.1: Describing uses of metals and metalloids** |
| 1. **(i) bronze; (ii) brass; (iii) solder; (iv) steel**

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| **Image result for test iconTest your knowledge 7.2: Describing uses of the hydroxides, oxides and salts of sodium and calcium** |
| 1. **Sodium chloride – flavouring food and de-icing roads; sodium nitrate – fertiliser; sodium sulphate - detergent**
2. **Calcium chloride – drying agent, making roads less dusty**
3. **NaOH: make soap, make paper; CaO – drying agent, remove SiO2 from iron; Ca(OH)2 – water treatment and cement production**
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***Lesson 8 – Why are metals and their compounds useful (part II)?***

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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 8.1: Form complex ions by reacting metal ions with excess ammonia** |
| Equipment needed: 0.1 moldm-3 solutions of any soluble salt of Pb2+, Ca2+, Fe2+, Fe3+, Zn2+, Al3+ and Cu2+ - one bottle of each is sufficient - each bottle should come with its own dropping pipette - 2 cm3 per group; 1 – 2 moldm-3 ammonia solution - one bottle per group - 100 cm3 per group needed; 7 test tubes and one test tube rack per groupExpected observations:Pb2+(aq) + 2OH-(aq) 🡪 Pb(OH)2(s); Ca2+(aq) + 2OH-(aq) 🡪 Ca(OH)2(s); Fe2+(aq) + 2OH-(aq) 🡪 Fe(OH)2(s); Fe3+(aq) + 3OH-(aq) 🡪 Fe(OH)3(s); Zn2+(aq) + 2OH-(aq) 🡪 Zn(OH)2(s); Al3+(aq) + 3OH-(aq) 🡪 Al(OH)3(s); Cu2+(aq) + 2OH-(aq) 🡪 Cu(OH)2(s)Zn(OH)2(s) + 6NH3(aq) 🡪 [Zn(NH3)6]2+(aq) + 2OH-(aq); Cu(OH)2(s) + 4NH3(aq) 🡪 [Cu(NH3)4]2+(aq) + 2OH-(aq)Pb(OH)2 and Al(OH)3 dissolve in excess NaOH but not excess NH3; Cu(OH)2 dissolves in excess NH3 but not excess NaOH |

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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 8.2: React anhydrous copper sulphate with water** |
| Chemicals needed: anhydrous CuSO4 (5 g per group); one bottle per class each of paraffin and ethanol, each with its own dropping pipetteApparatus needed per group: three watch glasses and one spatula, access to mass balanceThe water will turn anhydrous copper sulphate blue; the paraffin should not; the ethanol might turn the copper sulphate slightly blue if it also contains water |

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| **Image result for test iconTest your knowledge 8.3: Describing properties and reactions of compounds of d-block metals** |
| 1. **CuCl2 (catalyst); CuSO4 (fungicide)**
2. **CuO (pigment/disposal of toxic compounds)**
3. **Complex ion: species containing a central metal ion attached to one or more ligands by dative covalent bonds; ligand: species with a lone pair of electrons which it can use to form a dative bond with a metal ion; eg [Zn(NH)6]2+ or [Cu(NH3)4(H2O)2]2+**
4. **Electrons in partially filled d-orbitals of complex ions can absorb visible light**
5. **D-orbitals fully filled**
6. **No complex ion present**
7. **Cu2+(aq) + 4NH3 + 2H2O 🡪 [Cu(NH3)4(H2O)2]2+; pale blue precipitate appears, then a dark blue solution**
8. **Zn2+(aq) + 6NH3 🡪 [Zn(NH)6]2+; white precipitate appears, then a colourless solution**
9. **Add a few drops of the liquid to anhydrous copper sulphate; if a blue colour is formed, water is present;** CuSO4(s) + 5H2O(l) 🡪 CuSO4.5H2O(s)
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***Lesson 9 - How can we use complex formation reactions in qualitative analysis?***

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| **Summary Activity 9.1: What can you remember about qualitative analysis?** |
| * The experimental identification of a substance of species present in a substance
* Fe2+ (dark green), Fe3+ (orange); Ca2+(white); Al3+ (white); Pb2+ (white); Cu2+ (pale blue); Zn2+ (white)
* Al(OH)3, Pb(OH)2and Zn(OH)2
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| cid:ii_jepnvfe00_1621fa54a497745d **Practical 9.2: Use complex formation reactions to identify cations in solution** |
| Equipment needed: 0.1 moldm-3 solutions of any soluble salt of Pb2+, Ca2+, Zn2+ and Al3+ - one bottle of each is sufficient; they should be labelled A, B, C and D - each bottle should come with its own dropping pipette - 5 cm3 per group; 1 – 2 moldm-3 ammonia solution - one bottle per group - 50 cm3 per group needed; 0.5 - 1 moldm-3 NaOH solution – one bottle per group - 50 cm3 per group needed; 8 test tubes and one test tube rack per groupExpected observations and results:Pb2+ and Al3+ cannot be distinguished by these tests; Pb2+ gives a precipitate with Cl- ions but Al3+ does not, so the addition of a few drops of hydrochloric acid will give a white precipitate with the solution containing Pb2+ but not the solution containing Al3+ |

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| Image result for test icon**Test your knowledge 9.3: Understanding qualitative analysis by complex formation reactions** |
| Answer: add dilute ammonia dropwise until in excess; both solutions will give a white precipitate; the precipitate formed from the solution of zinc sulphate will dissolve in excess ammonia but the precipitate formed from the solution of aluminium sulphate will not |

***Lesson 10 – What have I learned about metals and their compounds?***

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| Image result for test icon**10.1 END-OF-UNIT QUIZ****UNIT 9 – METALS AND THEIR COMPOUNDS** |
| 1. Lattice of cations, held together by a sea of delocalised electrons; cations can move past each other without breaking the attraction between cations and electrons
2. Electrical conductors (delocalised electrons); sonorous (sound waves can travel through with little loss of energy); lustrous (electrons reflect light back to its source)
3. (a) Mg + 2HNO3 🡪 Mg(NO3)2 + H2; (b) 2Na + 2H2O 🡪 2NaOH + H2; (c) Zn + CuSO 🡪 ZnSO4 + Cu; redox reactions
4. Zinc is more reactive than copper so zinc can displace copper from its compounds; copper is less reactive than zinc so copper cannot displace zinc from its compounds
5. Purified Al2O3 is dissolved in molten cryolite and electrolysed using graphite anodes; molten aluminium is produced at the cathode
6. Brass – used in taps due to its anti-bacterial properties; made from copper and zinc; solder is used to weld electrical components together due to its low melting point; made from tin and lead
7. Hydrated Cu2+ ions have a d9 configuration so can absorb visible light; hydrated Zn2+ ions have a d10 configuration so cannot
8. d-block metals have s and d electrons but no p-electrons in their outer shell; transition metals can form at least one stable ion with partially filled d-orbitals; all transition metals come from the d-block but not all d-block metals are transition metals
9. the number of d-electrons lost by transition metals can vary depending on the reaction; the energy required to remove the d-electrons is sometimes but not always recovered in bonding
10. with Zn2+; white precipitate, which dissolves in excess ammonia to give a colourless solution: Zn2+(aq) + 2OH-(aq) 🡪 Zn(OH)2(s); Zn(OH)2(s) + 6NH3(aq) 🡪 [Zn(NH3)6]2+(aq) + 2OH-(aq); with Cu2+; pale blue precipitate, which dissolves in excess ammonia to give a deep blue solution: Cu2+(aq) + 2OH-(aq) 🡪 Cu(OH)2(s); Cu(OH)2(s) + 4NH3(aq) + 2H2O(l) 🡪 [Zn(NH3)4(H2O)2]2+(aq) + 2OH-(aq)
11. add aqueous NH3 to both gradually until in excess; with Zn2+, a white precipitate will form which dissolves in excess NaOH; with Al3+, a white precipitate will form which is insoluble in excess NaOH
12. Add a few drops of the liquid to anhydrous copper sulphate; if it turns blue, water is present.
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