UNIT 9

METALS AND THEIR COMPOUNDS

Answers

Lesson 1 – What are metals?

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)

Test your knowledge 1.2: Classifying metals, non-metals and metalloids

(a) Eg sodium, calcium, magnesium, potassium
(b) Eg aluminium, tin, lead
(c) Eg copper, iron, zinc
(d) Eg boron, silicon
(e) Eg oxygen, bromine, neon
(f) Electronegativity increases across a Period, so the attraction to bonding electrons increases, so atoms become less likely to allow bonding electrons to delocalise
(g) 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?

Test your knowledge 2.1: Describing physical properties of metals

(a) Delocalised electrons are free to move
(b) cations can move around without breaking metallic bonds
(c) Mg^{2+} 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
(d) 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
(e) Al^{3+} is smaller than Mg^{2+} and is more highly charged, so it attracts delocalised electrons more strongly, so the metallic bonding is stronger
(f) Iron has a larger atomic mass than aluminium
(g) Iron has more unpaired electrons than copper

Lesson 3 – How do s and p-block metals react with air, water and acids?

Summary Activity 3.1: What can you remember about redox reactions of metals?

- Eg $2Mg + O_2 \rightarrow 2MgO$ or $4Al + 3O_2 \rightarrow 2Al_2O_3$ (metal oxidised, O reduced)
- Eg $Mg + 2H^+ \rightarrow Mg^{2+} + H_2$ or $Zn + 2H^+ \rightarrow Zn^{2+} + H_2$ (metal oxidised, H^+ reduced)
- Eg $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ (Zn oxidised, Cu^{2+} reduced)

Test your knowledge 3.2: Describing chemical properties of s and p-block metals

(a) $2K + 2H_2O \rightarrow 2KOH + H_2$; $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$; 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
(b) $2Na + O_2 \rightarrow Na_2O_2$; $2Ca + O_2 \rightarrow 2CaO$; sodium forms a peroxide, calcium forms an oxide
(c) To prevent them from reacting with air or water
(d) $Mg + 2HCl \rightarrow MgCl_2 + H_2$ or $Mg + 2H^+ \rightarrow Mg^{2+} + H_2$; $2Al + 6HCl \rightarrow AlCl_3 + 3H_2$ or $2Al + 6H^+ \rightarrow 2Al^{3+} + 3H_2$; 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
(e) Al forms a very stable oxide layer on its surface which protects it from further reaction

Lesson 4 - How do d-block metals react with air, water and acids?

Summary Activity 4.1: What can you remember about d-block metals?

- Iron rusts when is is oxidised by O_2 and H_2O 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 $CuSO_4$ using copper electrodes; the copper on the impure anode dissolves ($Cu \rightarrow Cu^{2+} + 2e$ and pure copper is deposited at the cathode: $Cu^{2+} + 2e \rightarrow Cu$
- Cu: $1s^22s^22p^63s^23p^64s^13d^{10}$; Fe: $1s^22s^22p^63s^23p^64s^23d^6$
- Cu^{2+}: $1s^22s^22p^63s^23p^63d^9$; Fe^{2+}: $1s^22s^22p^63s^23p^63d^6$; Fe^{3+}: $1s^22s^22p^63s^23p^63d^5$

Test your knowledge 4.2: Describing chemical properties of d-block metals

(a) **because they do not always lose all of their valence d-electrons; the number of electrons they lose can vary**
(b) **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)**
(c) **they can change their oxidation state so can accept and donate electrons; eg Fe in Haber process**
(d) $4Fe + 6H_2O + 3O_2 \rightarrow 4Fe(OH)_3$
(e) **Copper does not react with water; gold does not react with air or water so stays shiny for a long time**
(f) **$Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$**

Lesson 5 – How can we compare the reactivity of different metals?

Summary Activity 5.1: Which types of redox reaction involve metals?

- $2Na + 2H_2O \rightarrow 2NaOH + H_2$; Na oxidised from 0 to +1; H reduced from +1 to 0; zinc less reactive than sodium and cannot displace hydrogen from water
- $Zn + 2HCl \rightarrow ZnCl + H_2$; Zn oxidised from 0 to +2; H reduced from +1 to 0; copper less reactive than zinc and cannot displace hydrogen from acids
- $Zn + CuSO_4 \rightarrow Cu + ZnSO_4$; Zn oxidised from 0 to +2; Cu reduced from +2 to 0; copper less reactive than zinc so cannot displace zinc from its compounds
- $ZnO + C \rightarrow Zn + CO$; 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
- $CuO + H_2 \rightarrow Cu + H_2O$; 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

Practical 5.2: Compare the reactivity of different metals

Chemicals needed per group: 12 test tubes, three 10 cm^3 measurin cylinders, one test tube rack, one thermometer access to 0.5 $moldm^{-3}$ solutions of $CuSO_4$, $ZnSO_4$, $MgSO_4$ and $FeSO_4$ (5 cm^3 per group), each bottle with its own dropping pipette; access to powdered samples of Zn, Fe, Cu and Mg, each with its own spatula

Signs 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 $CuSO_4$; the reactions of Zn with $CuSO_4$ and Mg with $FeSO_4$ may also be vigorous.

Metal powder	salt solution			
	$ZnSO_4$	$CuSO_4$	$FeSO_4$	$MgSO_4$
Zn		vigorous reaction – large temperature change, orange solid produced	Reaction – small temperature change	No visible reaction /temperature change
Cu	No visible reaction /temperature change		No visible reaction /temperature change	No visible reaction /temperature change
Fe	No visible reaction /temperature change	Reaction – small temperature change, orange solid produced		No visible reaction /temperature change
Mg	Reaction – small temperature change	Very vigorous reaction – very large temperature change, orange solid produced	vigorous reaction – large temperature change	

$CuSO_4$ with Mg, Zn and Fe: $CuSO_4 + Mg \rightarrow MgSO_4 + Cu$; $CuSO_4 + Zn \rightarrow ZnSO_4 + Cu$; $CuSO_4 + Fe \rightarrow FeSO_4 + Cu$

$FeSO_4$ with Mg and Zn: $FeSO_4 + Mg \rightarrow MgSO_4 + Fe$; $FeSO_4 + Zn \rightarrow ZnSO_4 + Fe$

$ZnSO_4$ with Mg: $ZnSO_4 + Mg \rightarrow MgSO_4 + Zn$

Mg 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

Test your knowledge 5.3: Understanding metal displacement reactions

(a) no reaction; (b) $Mg + CuSO_4 \rightarrow MgSO_4 + Cu$; (c) no reaction; (d) no reaction; (e) $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$; (f) no reaction; (g) $Fe_2O_3 + 3C \rightarrow Fe_2O_3 + 3CO$; (h) $SnO_2 + 2C \rightarrow Sn + 2CO$

Lesson 6 - How are metals extracted from their ores?

Summary Activity 6.1: Electrolytic Processes

- Cathode: $Al^{3+} + 3e^- \rightarrow Al$; anode $2O^{2-} \rightarrow O_2 + 4e^-$
- Cathode: $Cu^{2+} + 2e \rightarrow Cu$; anode: $Cu \rightarrow Cu^{2+} + 2e^-$

Test your knowledge 6.2: Understanding the Extraction of Metals

(a) Reactivity, required purity, cost of process
(b) Lots of energy needed to melt the cryolite and for the electricity
(c) The melting point of cryolite is lower than the melting point of pure aluminium oxide
(d) $Al^{3+} + 3e \rightarrow Al$ (at the cathode) and $2O^{2-} \rightarrow O_2 + 4e$ (at the anode)
(e) Anodes react with oxygen $C + O_2 \rightarrow CO_2$
(f) $C + O_2 \rightarrow CO_2$; $C + CO_2 \rightarrow 2CO$
(g) $SnO_2 + 2CO \rightarrow Sn + 2CO_2$
(h) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
(i) It helps remove the main impurity SiO_2; $CaCO_3$ decomposes to produce CaO, which reacts with SiO_2 to produce $CaSiO_3$, which can be removed
(j) Oxygen is bubbled through the molten iron; the oxygen removes the C as CO_2: $C + O_2 \rightarrow CO_2$
(k) Al is more reactive than C and cannot be reduced from its oxide by C or CO
(l) $CaSiO_3$ is used in road-building
(m) Fe is magnetic so can be separated from other scrap using a magnet
(n) It is present in low concentrations and is difficult to obtain in pure form
(o) 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

Lesson 7 – Why are metals and their compounds useful (part I)?

Test your knowledge 7.1: Describing uses of metals and metalloids

(a) (i) bronze; (ii) brass; (iii) solder; (iv) steel
(b)

Material	Use	Property
aluminium	Aircraft	Low density, strong, resistant to corrosion
tin	Prevent corrosion of iron	Forms stable oxide layer
solder	Welding metal parts together	Low melting point
steel	construction	Strong and cheap
Gold	jewellery	Unreactive so stays shiny
copper	Water pipes	Doesn't react with water
brass	taps	Has anti-bacterial properties
silicon	Electronic components	Semiconductor

Test your knowledge 7.2: Describing uses of the hydroxides, oxides and salts of sodium and calcium

(a) Sodium chloride – flavouring food and de-icing roads; sodium nitrate – fertiliser; sodium sulphate - detergent
(b) Calcium chloride – drying agent, making roads less dusty
(c) NaOH: make soap, make paper; CaO – drying agent, remove SiO_2 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)?

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 Pb^{2+}, Ca^{2+}, Fe^{2+}, Fe^{3+}, Zn^{2+}, Al^{3+} and Cu^{2+} - one bottle of each is sufficient - each bottle should come with its own dropping pipette - 2 cm^3 per group; 1 – 2 $moldm^{-3}$ ammonia solution - one bottle per group - 100 cm^3 per group needed; 7 test tubes and one test tube rack per group

Expected observations:

Cation present in solution	Observation on adding a few drops of ammonia	Observation on adding excess ammonia
Pb^{2+}	white precipitate	no change
Ca^{2+}	white precipitate	no change
Fe^{2+}	dark green precipitate	no change
Fe^{3+}	orange/brown precipitate	no change
Zn^{2+}	white precipitate	precipitate dissolves; colourless solution formed
Al^{3+}	white precipitate	no change
Cu^{2+}	pale blue precipitate	precipitate dissolves; dark blue solution formed

$Pb^{2+}(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(s)$; $Ca^{2+}(aq) + 2OH^-(aq) \rightarrow Ca(OH)_2(s)$; $Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$; $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$; $Zn^{2+}(aq) + 2OH^-(aq) \rightarrow Zn(OH)_2(s)$; $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$; $Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$

$Zn(OH)_2(s) + 6NH_3(aq) \rightarrow [Zn(NH_3)_6]^{2+}(aq) + 2OH^-(aq)$; $Cu(OH)_2(s) + 4NH_3(aq) \rightarrow [Cu(NH_3)_4]^{2+}(aq) + 2OH^-(aq)$

$Pb(OH)_2$ and $Al(OH)_3$ dissolve in excess NaOH but not excess NH_3; $Cu(OH)_2$ dissolves in excess NH_3 but not excess NaOH

Practical 8.2: React anhydrous copper sulphate with water

Chemicals needed: anhydrous $CuSO_4$ (5 g per group); one bottle per class each of paraffin and ethanol, each with its own dropping pipette

Apparatus needed per group: three watch glasses and one spatula, access to mass balance

The 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

Test your knowledge 8.3: Describing properties and reactions of compounds of d-block metals

(a) **$CuCl_2$ (catalyst); $CuSO_4$ (fungicide)**

(b) **CuO (pigment/disposal of toxic compounds)**

(c) **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 $[Cu(NH_3)_4(H_2O)_2]^{2+}$**

(d) **Electrons in partially filled d-orbitals of complex ions can absorb visible light**

(e) **D-orbitals fully filled**

(f) **No complex ion present**

(g) **$Cu^{2+}(aq) + 4NH_3 + 2H_2O \rightarrow [Cu(NH_3)_4(H_2O)_2]^{2+}$; pale blue precipitate appears, then a dark blue solution**

(h) **$Zn^{2+}(aq) + 6NH_3 \rightarrow [Zn(NH)_6]^{2+}$; white precipitate appears, then a colourless solution**

(i) **Add a few drops of the liquid to anhydrous copper sulphate; if a blue colour is formed, water is present;** $CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4.5H_2O(s)$

UNIT 9 – METALS AND THEIR COMPOUNDS

Lesson 9 - How can we use complex formation reactions in qualitative analysis?

Summary Activity 9.1: What can you remember about qualitative analysis?

- The experimental identification of a substance of species present in a substance
- Fe^{2+} (dark green), Fe^{3+} (orange); Ca^{2+}(white); Al^{3+} (white); Pb^{2+} (white); Cu^{2+} (pale blue); Zn^{2+} (white)
- $Al(OH)_3$, $Pb(OH)_2$ and $Zn(OH)_2$

Practical 9.2: Use complex formation reactions to identify cations in solution

Equipment needed: 0.1 $moldm^{-3}$ solutions of any soluble salt of Pb^{2+}, Ca^{2+}, Zn^{2+} and Al^{3+} - 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 cm^3 per group; 1 – 2 $moldm^{-3}$ ammonia solution - one bottle per group - 50 cm^3 per group needed; 0.5 - 1 $moldm^{-3}$ NaOH solution – one bottle per group - 50 cm^3 per group needed; 8 test tubes and one test tube rack per group

Expected observations and results:

Solution	Observations				Cation Present
	Few drops NaOH	Excess NaOH	Few drops NH_3	Excess NH_3	
A	white precipitate	dissolves - colourless solution	white precipitate	no change	Pb^{2+} or Al^{3+}
B	white precipitate	no change	white precipitate	no change	Ca^{2+}
C	white precipitate	dissolves - colourless solution	white precipitate	dissolves - colourless solution	Zn^{2+}
D	white precipitate	dissolves - colourless solution	white precipitate	no change	Pb^{2+} or Al^{3+}

Pb^{2+} and Al^{3+} cannot be distinguished by these tests; Pb^{2+} gives a precipitate with Cl^- ions but Al^{3+} does not, so the addition of a few drops of hydrochloric acid will give a white precipitate with the solution containing Pb^{2+} but not the solution containing Al^{3+}

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

UNIT 9 – METALS AND THEIR COMPOUNDS

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

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) $Mg + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2$; (b) $2Na + 2H_2O \rightarrow 2NaOH + H_2$; (c) $Zn + CuSO \rightarrow ZnSO_4 + Cu$; 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 Al_2O_3 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 Cu^{2+} ions have a d^9 configuration so can absorb visible light; hydrated Zn^{2+} ions have a d^{10} 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 Zn^{2+}; white precipitate, which dissolves in excess ammonia to give a colourless solution: $Zn^{2+}(aq) + 2OH^-(aq) \rightarrow Zn(OH)_2(s)$; $Zn(OH)_2(s) + 6NH_3(aq) \rightarrow [Zn(NH_3)_6]^{2+}(aq) + 2OH^-(aq)$; with Cu^{2+}; pale blue precipitate, which dissolves in excess ammonia to give a deep blue solution: $Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$; $Cu(OH)_2(s) + 4NH_3(aq) + 2H_2O(l) \rightarrow [Zn(NH_3)_4(H_2O)_2]^{2+}(aq) + 2OH^-(aq)$
11. add aqueous NH_3 to both gradually until in excess; with Zn^{2+}, a white precipitate will form which dissolves in excess NaOH; with Al^{3+}, 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.