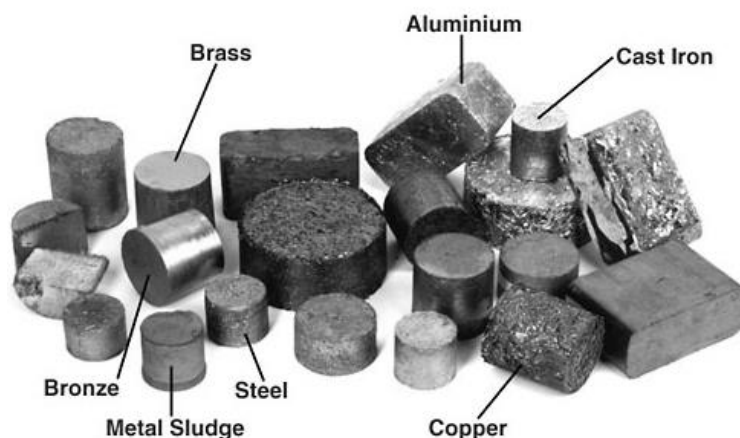


# UNIT 9

## METALS AND THEIR COMPOUNDS

### Student Version



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### Contents

- a) Classification – Metals, Non-metals and Metalloids
- b) Physical Properties of Metals
- c) Chemical Properties of Metals
- d) Extraction of Metals
- e) Uses of Metals and Metalloids
- f) Properties and Reactions of Metal Compounds
- g) Qualitative Analysis Part 4

Key words: metal, non-metal, metalloid, malleability, ductility, lustre, sonority, alloy, bronze, brass, steel, solder, malleable, ductile, sonorous, lustrous, ferromagnetic, s-block metal, p-block metal, d-block metal, transition metal, rusting, reactivity series, metal displacement reaction, blast furnace, basic oxygen process, complex ion, ligand

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

- Unit 1 – Atoms and the Periodic Table
- Unit 2 – Particles, Bonding and Structure
- 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

**Estimated Teaching Time: 10 hours**

Lesson 1 – What are metals?

## a) Classification – Metals, Non-metals and Metalloids



### Summary Activity 1.1: What can you remember about metals?

- What is a metal?
- Describe the structure and bonding in a typical metal
- What are the properties of a typical metal?
- What is a non-metal?
- Describe the structure and bonding in a typical non-metal
- Why are some elements metals and others non-metals?
- What is an alloy? How many alloys can you name?

#### (i) Metals

- **Metals** are elements which contain metallic bonding
- In metallic bonding, the atoms lose some or all of their outer shell (valence) electrons to form a lattice of cations and a sea of delocalised electrons; the delocalised electrons and the cations are attracted to each other and this holds the structure together
- Metallic bonding is preferred when all of the bonding atoms are electropositive, and hence do not hold on tightly to their valence electrons; it is therefore most commonly found in the elements of atoms on the left-hand side, and at the bottom, of the Periodic Table (bonding in elements becomes less metallic and more covalent from left to right across a Period, and more metallic and less covalent from top to bottom in a Group)
- In some cases, two or more different atoms can form a metallic bond; usually if both atoms are electropositive (ie come from metals; eg Cu and Zn); electropositive atoms can also form metallic bonds with atoms which are not electropositive (ie come from non-metals), provided that the electropositive atom is more abundant (eg Fe and C); mixtures which contain metallic bonds are known as **alloys**; alloys are not compounds because they do not have a fixed composition

#### (ii) Non-metals

- **Non-metals** are elements which do not contain metallic bonds; the vast majority of non-metals contain covalent bonding
- In elements, covalent bonding is preferred to metallic bonding if the atoms are not electropositive enough to give up their valence electrons; it is therefore most commonly found in the elements of atoms on the right-hand side, and at the top of the Periodic Table (bonding in elements becomes less metallic and more covalent from left to right across a Period, and more metallic and less covalent from top to bottom in a Group); elements containing covalent bonds are called **non-metals**

## Unit 9 – Metals and their Compounds

### (iii) Metalloids

- **Metalloids** or **semi-metals** are substances which contain bonding intermediate in character between metallic and covalent; metalloids can be elements, compounds or mixtures, but the term is most commonly used to describe elements; often these substances display metallic properties at higher temperatures and non-metallic properties at lower temperatures; they are found in between the metals and the non-metals in the Periodic Table
- The following diagram shows the location of the metals, non-metals and metalloids in the Periodic Table; note the trend from metal to metalloid to non-metal across a Period from right to left and up a Group from bottom to top: (white = metal, grey = metalloid; black = non-metal)

I	II												III	IV	V	VI	VII	0	
																			He
Li	Be												B	C	N	O	F	Ne	
Na	Mg												Al	Si	P	S	Cl	Ar	
K	Ca	Sc											Ga	Ge	As	Se	Br	Kr	
Rb	Sr	Y											In	Sn	Sb	Te	I	Xe	
Cs	Ba	La	Ce - Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	
Fr	Ra	Ac	Th - Lw																

### (iv) Further classification of metals

- Metals can be further classified into s-block, p-block, d-block or f-block metals based on their electronic configuration and consequent position in the Periodic Table:
- Metals made from atoms containing valence electrons in s-orbitals only are called s-block metals (eg Na, Ca)
- Metals made from atoms containing valence electrons in s and d-orbitals only are called d-block metals (eg Fe, Cu, Au)
- Metals made from atoms containing valence electrons in s and p-orbitals only are called p-block metals (eg Al)



#### Test your knowledge 1.2: Metals, non-metals and metalloids

- Name two s-block metals
- Name two p-block metals
- Name two d-block metals
- Name two metalloids
- Name two non-metals
- Why does metallic character decrease across a Period?
- Why does metallic character increase down a Group?

Lesson 2 – What are the physical properties of metals?

## b) Physical Properties of Metals

- Most physical properties of metals are caused by the metallic bonding; in most cases, variations in the physical properties of the metals are caused by variations in the strength of metallic bonding:

### (i) Strength of Metallic Bonding

- The strength of metallic bonding depends on:
  - The charge on the cations: the higher the charge on the cations, the stronger the metallic bonding  
Eg  $\text{Mg}^{2+}$  is smaller than  $\text{Na}^+$  so the metallic bonding is stronger in Mg than in Na
  - The size of the cations: the smaller the size of the cations, the stronger the metallic bonding  
Eg  $\text{Na}^+$  is smaller than  $\text{K}^+$  so the metallic bonding is stronger in Na than in K
  - The number of delocalised electrons per atom: the greater the number of delocalised electrons, the stronger the metallic bonding; the number of delocalised electrons per atom is equal to the charge on the cation  
Eg Mg gives up two electrons, Na gives up only one, so the metallic bonding is stronger in Mg than in Na
  - The block of the Periodic Table; the bonding in d-block metals is much stronger than the bonding in s and p-block metals (the reasons for this will not be discussed in this course)

### (ii) Physical Properties of Metals

- Some physical properties are common to all metals:
  - **Electrical conductivity:** all metals contain delocalised electrons, which can move when a potential difference is applied; all metals are therefore good conductors of electricity
  - **Thermal conductivity:** the delocalised electrons carry heat energy with them, which can therefore move rapidly through the structure as the electrons move; all metals are therefore good conductors of heat
  - **Malleability:** in metals, the attraction between the delocalised electrons and the cations does not depend on the position of the cations; it is therefore possible to move the cations around without having a significant impact on the bonding; this means that metals can be easily bent into different shapes; all metals are therefore malleable
  - **Ductility:** because metals can be bent into different shapes, they can also be turned into wires; this property is known as ductility; all metals are therefore ductile, especially those with stronger metallic bonding
  - **Lustre:** the delocalised electrons in metals reflect light back towards its source; this causes metals to have a characteristic shiny appearance known as lustre
  - **Sonority:** the delocalised electrons in metals can also transport sound energy efficiently through the solid, which means that the sound wave can travel through the metal with little loss of energy; the ability of a solid to make an object when hit is known as sonority; most metals are sonorous (although high density metals are less sonorous)

## Unit 9 – Metals and their Compounds

- Some other physical properties of metals depend on the strength of the metallic bonding, which varies significantly from metal to metal:
  - **Melting point:** the melting point of metals depends on the strength of the metallic bonding; it is generally high as metallic bonding is generally strong, so a lot of energy is needed to break up the metallic lattice; however the strength of the metallic bonding varies widely between metals and so does the melting point. Group I metals have weak metallic bonding and a low melting point; d-block metals have strong metallic bonding and a high melting point
  - **Tensile strength:** because metals are malleable, they can absorb large quantities of energy without break and so tend to have a high tensile strength; however the tensile strength depends on the strength of the metallic bonding, which varies widely between metals; Group I metals have weak metallic bonding and a tensile strength; d-block metals have strong metallic bonding and a high tensile strength
  
- Some properties of metals do not result from either the delocalised electrons or the strength of the metallic bonding:
  - **Magnetic Properties:** some metals have their delocalised electrons arranged in such a way that they can align themselves to a magnetic field and attract a magnet; the ability to do this is known as **ferromagnetism**; only Fe, Co and Ni are ferromagnetic; ferromagnetism is caused by the presence of unpaired electrons in the atoms: Fe, Co and Ni have electronic configurations  $d^6$ ,  $d^7$  and  $d^8$  respectively, which means that they have four, three and two unpaired electrons respectively; however not all atoms with unpaired electrons are ferromagnetic and the full causes of ferromagnetism will not be discussed in this course)
  - **Density:** density is mass per unit volume; the density of a metal therefore depends on the atomic mass (the higher the atomic mass the higher the density) and the atomic size (the larger the atomic size the lower the density); the density also depends to an extent on the type of crystal structure adopted by the metal, as some crystal structures are more compact than others; atomic mass increases across a Period and atomic size decrease across a period; the density of metals therefore increases across a Period; atomic mass increases down a group but so does atomic size; generally, the increase in atomic mass down a group is more significant so density tends to increase down a group; the metallic elements with the highest density are therefore to be found at the bottom right of the Periodic Table
  
- **Summary of properties of metals:**

Property	s-block and p-block	d-block
Electrical and Thermal Conductivity	High	high
Melting Point	low-medium	high
Tensile strength	low-medium	high
Malleability	High	high
Ductility	Medium	high
Lustre	High	high
Sonority	High	High
Magnetism	not-magnetic	Fe, Co, Ni magnetic
Density	low-medium	medium-high



**Test your knowledge 2.1: Physical properties of metals**

- (a) Why do all metals conduct electricity?
- (b) Why are all metals malleable?
- (c) Why does magnesium have a higher melting point than sodium?
- (d) Why does potassium have a lower melting point than potassium?
- (e) Why is aluminium stronger than magnesium?
- (f) Why is iron more dense than aluminium?
- (g) Why is iron more magnetic 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?**

- Give one example of a reaction between a metal and oxygen
- Give one example of a reaction between a metal and an acid
- Give one example of a metal displacement reaction
- In each reaction, what is oxidised and what is reduced?

## c) Chemical Properties of Metals

**(i) s-block metals (Na and Ca)**

- s-block metals are electropositive and do not hold on to their electrons strongly; the valence electrons are easily lost during chemical reactions and the metal is oxidised; most s-block metals react rapidly with air, water and acids as a result
- s-block metals only form one stable cation, corresponding to the loss of all valence electrons; Group 1 metals form +1 ions only (eg  $\text{Na}^+$ ) and Group 2 metals form +2 ions only (eg  $\text{Ca}^{2+}$ )
- Chemical reactivity increases as electropositivity increases; reactivity therefore decreases down a Group and Group 1 metals are more reactive than Group 2 metals
- With air, the elements will react to form oxides (containing  $\text{O}^{2-}$ ) (eg  $\text{CaO}$ )
  - although barium and all Group 1 metals except lithium can also form peroxides (containing  $\text{O}_2^{2-}$ ) (eg  $\text{Na}_2\text{O}_2$ )
  - these reactions are violent and exothermic once the metal is ignited
  - the more reactive s-block metals can ignite spontaneously, most s-block metals tarnish rapidly on exposure to air (they form an oxide layer at the surface of the metal which makes it less shiny)
  - $4\text{Na(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{Na}_2\text{O(s)}$ ;  $2\text{Na(s)} + \text{O}_2\text{(g)} \rightarrow \text{Na}_2\text{O}_2\text{(s)}$ ;  $2\text{Ca(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{CaO(s)}$
  - these reactions are redox reactions; the metal is oxidised and the oxygen is reduced; most s-block metals are stored under oil to prevent this reaction from taking place

## Unit 9 – Metals and their Compounds

- With water, the elements will react to form hydroxides (containing OH<sup>-</sup>) and hydrogen (H<sub>2</sub>)
  - all s-block elements except Be and Mg react rapidly with water
  - with the most reactive metals the reaction is very explosive; this is another reason why the s-block metals are stored under oil; the metal fizzes rapidly as hydrogen is given off
$$2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + \text{H}_2\text{(g)}; \text{Ca(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{Ca(OH)}_2\text{(aq)} + \text{H}_2\text{(g)}$$
  - these reactions are redox reactions; the metal is oxidised and the hydrogen in the water is reduced
  - watch these reactions on: [www.youtube.com/watch?v=fAFrzlyrqCU](http://www.youtube.com/watch?v=fAFrzlyrqCU)
  - with calcium, the solution quickly becomes cloudy; Ca(OH)<sub>2</sub> is only sparingly soluble in water and so the solution becomes saturated quickly; any more Ca(OH)<sub>2</sub> formed cannot dissolve and will form as a solid
- With acids, the reaction is even more vigorous and is dangerously explosive, especially with the Group 1 metals; the metal is oxidised and the H<sup>+</sup> ion in the acid is reduced (to hydrogen gas); a salt is also produced (see Topic 6 – Redox Reactions)
$$2\text{Na(s)} + 2\text{HCl(aq)} \rightarrow 2\text{NaCl(aq)} + \text{H}_2\text{(g)}; \text{Ca(s)} + 2\text{HCl(aq)} \rightarrow \text{CaCl}_2\text{(aq)} + \text{H}_2\text{(g)}$$

### (ii) p-block metals (Al, Sn and Pb)

- p-block metals are much less reactive than s-block metals; they are less electropositive and hold onto their electrons more strongly than s-block elements; of the common p-block metals, Al is the most reactive, followed by Sn and then lead
- aluminium only forms Al<sup>3+</sup> ions, corresponding to the loss of all valence electrons; tin can lose all of its electrons to form Sn<sup>4+</sup>; tin can also lose its p-electrons only (not its s-electrons) to form Sn<sup>2+</sup>; lead generally loses its p-electrons only to form Pb<sup>2+</sup>
- Although Al is considered a reactive metal, it shows almost no reaction with air or water; the reaction is much slower than expected because a small quantity of aluminium at the surface can react to form Al<sub>2</sub>O<sub>3</sub>; this oxide forms a very stable layer on the surface of the metal which prevents any further oxygen or water from reacting with the aluminium
- Aluminium powder can react with air if heated, and aluminium can react with steam if heated strongly:
$$4\text{Al(s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{Al}_2\text{O}_3\text{(s)}$$
$$4\text{Al(s)} + 6\text{H}_2\text{O(g)} \rightarrow 2\text{Al}_2\text{O}_3\text{(s)} + 3\text{H}_2\text{(g)}$$
The aluminium is oxidised; with oxygen, the O is reduced; with water, the H is reduced
- Most p-block metals, including Al, will react slowly with acids to give a salt and hydrogen:
$$2\text{Al(s)} + 6\text{HCl(aq)} \rightarrow 2\text{AlCl}_3\text{(aq)} + 3\text{H}_2\text{(g)}$$
The Al is oxidised and the H<sup>+</sup> in the acid is reduced (see Topic 6 – Redox Reactions)

## Unit 9 – Metals and their Compounds



### Test your knowledge 3.2: Chemical Properties of s and p-block metals

- Write equations for the reactions of potassium and calcium with water. Explain what you would observe, state the type of reaction taking place and explain why potassium reacts with water more quickly than calcium.
- Write equations for the reactions of sodium and calcium with oxygen. What is the difference between the two products?
- Why are the alkali metals stored under oil?
- Write equations for the reactions of magnesium and aluminium with hydrochloric acid. Explain the type of reaction taking place and explain why magnesium reacts with acids more quickly than aluminium.
- Explain why aluminium shows little reaction with air or water despite being considered a reactive metal.

### 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?

- Why does iron rust?
- How can rusting be prevented?
- How is copper purified?
- What are the electronic configurations of Cu and Fe?
- What are the electronic configurations of  $\text{Cu}^{2+}$ ,  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$ ?

#### (iii) d-block metals (Fe, Cu and Au)

- d-block metals are generally less reactive than p-block metals, although their electropositivity and reactivity does vary considerably; of the common d-block metals, Zn is the most reactive, followed by Fe; Cu is unreactive; Ag and Au are very unreactive
- d-block metals always lose their valence s-electrons; most d-block metals have two s-electrons so form 2+ ions (eg  $\text{Fe}^{2+}$ ) although atoms in columns 4 and 9 of the d-block only have one s-electron (eg Cu has the electronic configuration  $[\text{Ar}]3d^{10}4s^1$ ) so can form +1 ions (eg  $\text{Cu}^+$ ,  $\text{Au}^+$ ); the d-electrons in the outer shell, however, can behave as valence electrons in certain circumstances; this means that many atoms in the d-block can form several different ions, depending on how many d-electrons they lose:
  - Cu has the electronic configuration  $[\text{Ar}]3d^{10}4s^1$ ; it loses its valence s-electron to form  $\text{Cu}^+$ ; ( $[\text{Ar}]3d^{10}$ ); it usually also loses one valence d-electron to form  $\text{Cu}^{2+}$  ( $[\text{Ar}]3d^9$ )
  - Fe has the electronic configuration  $[\text{Ar}]3d^64s^2$ ; it loses its valence s-electrons to form  $\text{Fe}^{2+}$  ( $[\text{Ar}]3d^6$ ); it can also lose one valence d-electron to form  $\text{Fe}^{3+}$  ( $[\text{Ar}]3d^5$ )It is therefore common for d-block metals to form stable ions in which the d-orbitals are partially filled (ie ions which have 1 – 9 electrons in their d-orbitals); a metal which forms at least one stable ion with a partially filled d-orbital is called a **transition metal**; Cu, Fe and Au are all transition metals
- The ability of the d-electrons in transition metals to be lost and gained again results in transition metals (and their compounds) being good **catalysts**:
  - Fe is used as a catalyst in the Haber Process (the manufacture of ammonia:  $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$ )
  - Ni is used as a catalyst in the hydrogenation of alkenes ( $\text{C}_2\text{H}_4 + \text{H}_2 \rightleftharpoons \text{C}_2\text{H}_6$ )



## Unit 9 – Metals and their Compounds

- Reaction with air:
  - Most d-block metals show almost no reaction with air or water at normal temperatures, either because they are unreactive or because they form a stable oxide layer
  - Iron can react with air, when heated, to form iron (II) oxide and iron (III) oxide; copper can also react with air when heated to form copper (II) oxide;  $2\text{Fe(s)} + \text{O}_2 \rightarrow 2\text{FeO}$ ;  $4\text{Fe(s)} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$ ;  $2\text{Cu} + \text{O}_2 \rightarrow 2\text{CuO}$
  - Gold does not readily react with oxygen; this is why gold remains shiny for a long time
- Reaction with water:
  - Iron does not readily react with water in the absence of air but in the presence of air, iron can react with both oxygen and water to form rust:  $4\text{Fe} + 6\text{H}_2\text{O} + 3\text{O}_2 \rightarrow 4\text{Fe(OH)}_3$ ; unlike most other oxides and hydroxides of d-block elements, rust does not provide protection the iron surface and so iron will continue to rust if not protected
  - Copper is less reactive than iron and also forms a protective oxide layer; it does not react with water at all at room temperature but if it is exposed to hot water and air for a period of time it can form a green layer (copper carbonate) on its surface; this layer also provides protection to the copper
  - Gold does not readily react with water
- Reaction with acids:
  - Some d-block metals, including Fe and Zn will react slowly with acids to give a salt and hydrogen:  $\text{Fe(s)} + 2\text{HCl(aq)} \rightarrow \text{FeCl}_2\text{(aq)} + \text{H}_2\text{(g)}$ ; the Fe is oxidised and the  $\text{H}^+$  in the acid is reduced; this can be a useful way of preparing salts of d-block metals
  - Cu and Au do not react readily with dilute acids



### Test your knowledge 4.2: Chemical Properties of d-block metals

- Explain why d-block metals often have more than one stable oxidation state
- What is the difference between a d-block metal and a transition metal?
- Explain why transition metals are good catalysts and give an example of a reaction catalysed by a transition metal
- Write the equation for the rusting of iron
- Explain why copper is used in water pipes and why gold makes good jewellery
- Write an equation for the reaction between iron and sulphuric acid

## Unit 9 – Metals and their Compounds

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



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

- Write an equation to show how sodium reacts with water; indicate what is oxidised and what is reduced; why does zinc not react with water?
- Write an equation to show how zinc reacts with hydrochloric acid; indicate what is oxidised and what is reduced; why does copper not react with hydrochloric acid?
- Write an equation to show how zinc reacts with copper sulphate; indicate what is oxidised and what is reduced; why does copper not react with zinc sulphate?
- Write an equation to show how zinc oxide reacts with carbon; indicate what is oxidised and what is reduced; why does aluminium oxide not react with carbon?
- Write an equation to show how copper oxide reacts with hydrogen; indicate what is oxidised and what is reduced; why does zinc oxide not react with hydrogen?
- What is the general name given to all of these reactions?

#### (iv) The Reactivity Series of Metals

- The reactivity of different metals to oxygen, metals and water varies widely; it is possible to rank the reactivity of the different metals (and some non-metals); this is known as the reactivity series; the metals at the top are the most reactive:

Metal	Reactivity
Sodium	Reacts with cold water
calcium	Reacts with cold water
magnesium	Reacts with acids but not water
aluminium	Reacts with acids but not water
(carbon)	
zinc	Reacts with acids but not water
iron	Reacts with acids but not water
tin	Reacts with acids but not water
lead	Reacts with acids but not water
copper	No reaction with acids or water
gold	No reaction with acids or water

- The higher up the reactivity series a metal is, the more reactive it is and the more vigorously and quickly it will react with water, acids, oxygen and other non-metals; in addition, the more reactive the metal is, the harder it is to extract the metal from its compounds
- More reactive metals can displace less reactive metals from their compounds; for example, zinc will displace copper from copper sulphate solution:  $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$ ; these reactions are called **metal displacement** reactions; they also work in the solid state; for example, aluminium will reduce iron oxide to iron:  $2\text{Al} + \text{Fe}_2\text{O}_3 \rightarrow 2\text{Fe} + \text{Al}_2\text{O}_3$
- These reactions are all examples of redox reactions (see Topic 6 – Redox Reactions); the reactivity of the metals can be deduced directly from the standard electrode potentials, the most common of which are shown below:

## Unit 9 – Metals and their Compounds

Half-reaction	Conventional representation	E°/V
$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Mg}(\text{s})$	$\text{Mg} \text{Mg}^{2+}$	-2.37
$\text{Al}^{3+}(\text{aq}) + 3\text{e}^{-} \rightleftharpoons \text{Al}(\text{s})$	$\text{Al} \text{Al}^{3+}$	-1.66
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Zn}(\text{s})$	$\text{Zn} \text{Zn}^{2+}$	-0.76
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Fe}(\text{s})$	$\text{Fe} \text{Fe}^{2+}$	-0.44
$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Pb}(\text{s})$	$\text{Pb} \text{Pb}^{2+}$	-0.13
$2\text{H}^{+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{H}_2(\text{g})$	$\text{Pt} \text{H}_2, 2\text{H}^{+}$	0.00 (by definition)
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Cu}(\text{s})$	$\text{Cu} \text{Cu}^{2+}$	+0.34
$\text{Ag}^{+}(\text{aq}) + \text{e}^{-} \rightleftharpoons \text{Ag}(\text{s})$	$\text{Ag} \text{Ag}^{+}$	+0.80

- the more negative the standard electrode potential, the more reactive the metal; the above table shows clearly that the metals can be ranked in decreasing order of reactivity as follows: magnesium, aluminium, zinc, iron, lead, (hydrogen), copper, gold (see Unit 6 – Redox Reactions)



### Practical 5.2: Comparing the reactivity of different metals

- Using a measuring cylinder, pour 5 cm<sup>3</sup> of 0.5 moldm<sup>-3</sup> ZnSO<sub>4</sub> into each of three different test tubes
- Use a thermometer to measure the temperature of each solution
- Add a small spatula measure of magnesium powder to one test tube
- Observe what happens and note any temperature change
- Repeat steps (c) to (d) but this time adding copper powder instead of magnesium powder
- Repeat steps (c) to (d) but this time adding iron powder instead of magnesium powder
- Repeat steps (a) to (f) but this time use 5 cm<sup>3</sup> of 0.5 moldm<sup>-3</sup> CuSO<sub>4</sub>, and add zinc powder to the first test tube, magnesium powder to the second test-tube and iron powder to the third test tube
- Repeat steps (a) to (f) but this time use 5 cm<sup>3</sup> of 0.5 moldm<sup>-3</sup> FeSO<sub>4</sub>, and add zinc powder to the first test tube, magnesium powder to the second test-tube and copper powder to the third test tube
- Repeat steps (a) to (f) but this time use 5 cm<sup>3</sup> of 0.5 moldm<sup>-3</sup> MgSO<sub>4</sub>, and add zinc powder to the first test tube, copper powder to the second test-tube and iron powder to the third test tube
- Record your results in a table similar to the one below:

Metal powder	salt solution			
	ZnSO <sub>4</sub>	CuSO <sub>4</sub>	FeSO <sub>4</sub>	MgSO <sub>4</sub>
Zn				
Cu				
Fe				
Mg				

**In which mixtures does a reaction take place? Write an equation for each reaction.  
Can you list the four metals copper, magnesium, zinc and iron in order of reactivity?**



### Test your knowledge 5.3: metal displacement reactions

Predict whether a chemical reaction will take place when the following substances are mixed; write the symbol equation for any reaction occurring:

- (a) copper and magnesium sulphate
- (b) magnesium and copper sulphate
- (c) magnesium sulphate and zinc
- (d) zinc sulphate and copper
- (e) zinc and copper sulphate
- (f) aluminium oxide and carbon
- (g) iron (III) oxide and carbon
- (h) tin (IV) oxide and carbon

### Lesson 6 - How are metals extracted from their ores?



#### Summary Activity 6.1: Electrolytic Processes

- Give the half-equations for the reactions occurring at each electrode when Al is extracted from  $\text{Al}_2\text{O}_3$
- Give the half-equations for the reactions occurring at each electrode during the purification of copper

## (d) Extraction of Metals

### (i) General Principles of Extraction

- A mineral containing a sufficient quantity of a metal for its extraction to be commercially viable is called an **ore**
- Because metals are so useful, the extraction of metals from their ores is very commercially important; there are a number of different processes by which metals can be extracted from their ores; the choice of extraction process depends on:
  - the position of the metal in the reactivity series
  - the cost of the process
  - the required purity of the metal
- Metals can be extracted by electrolysis:
  - this works for all ores in which the bonding is ionic
  - it produces a pure form of the metal
  - but electrolysis is very expensive and only used for the most reactive metals
  - **aluminium**, for example, is extracted by the electrolysis of purified bauxite (aluminium oxide)
- Metals can also be extracted by reacting them with another element which is more reactive:
  - carbon is often used for this purpose as the process is very cheap
  - **iron, tin** and **copper** are extracted by reducing their ores (eg  $\text{Fe}_2\text{O}_3$  and  $\text{SnO}_2$ ) with carbon
  - this method cannot be used if the metal is above carbon in the reactivity series (like aluminium) or if the metal is not useful when contaminated with carbon (like titanium)
  - in such cases hydrogen or a reactive metal such as calcium should be used

## Unit 9 – Metals and their Compounds

- Metals which are present only in low concentrations require a more complex extraction process, and this is usually only commercially viable for very precious metals (like **gold**)

### (ii) Extraction of aluminium

- $\text{Al}_2\text{O}_3$  is purified and then electrolysed (see Unit 6 – Redox Reactions):
  - the  $\text{Al}_2\text{O}_3$  is dissolved in molten cryolite ( $\text{Na}_3\text{AlF}_6$ ) as it has a lower melting point than  $\text{Al}_2\text{O}_3$
  - the electrode reactions are:  $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$  (at the cathode) and  $2\text{O}^{2-} \rightarrow \text{O}_2 + 4\text{e}^-$  (at the anode)
  - graphite electrodes are used
  - the graphite anodes react with the oxygen produced and so they need to be regularly replaced
- This process uses a lot of energy and so is expensive; for this reason, recycling aluminium is very important

### (iii) Extraction of iron

- Iron is extracted from its main ore haematite ( $\text{Fe}_2\text{O}_3$ ) by reaction with carbon in a blast furnace; limestone is also added to remove the main impurity ( $\text{SiO}_2$ )
  - The carbon is added in the form of coke and it reacts as follows:  $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$ ;  $\text{C} + \text{CO}_2 \rightarrow 2\text{CO}$
  - the CO reduces the iron oxide to iron:  $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$
  - Some  $\text{Fe}_2\text{O}_3$  may be reduced by C directly:  $\text{Fe}_2\text{O}_3 + 3\text{C} \rightarrow 2\text{Fe} + 3\text{CO}$
  - limestone ( $\text{CaCO}_3$ ) is also added; it decomposes on heating to give CaO:  $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
  - the CaO (which is basic) reacts with the main impurity  $\text{SiO}_2$  (which is acidic):  $\text{CaO} + \text{SiO}_2 \rightarrow \text{CaSiO}_3$
  - the  $\text{CaSiO}_3$  is known as “slag”; it has a low density and can be separated from the reaction mixture;  $\text{CaSiO}_3$  is used in road surfacing; it is laid underneath the tar
- The iron produced by this method contains a large quantity of carbon (up to 4%); some carbon is desirable (pure iron combined with a small quantity of iron makes the alloy steel) but 4% is too high; it is therefore necessary to blow oxygen through the molten iron in order to remove some of the C as CO and  $\text{CO}_2$ ; other impurities such as sulphur can also be removed in this way; this is called the **basic oxygen process**
- the amount of carbon to be removed depends on the amount of carbon wanted in the steel; low-carbon steel is soft and used to make decorative items such as railings; medium-carbon steel is harder but still flexible and mainly used in construction; high-carbon steel is hard and brittle and used in tools
- iron is easy to recycle; it can be separated from other metals due to its magnetic properties

### (iv) Extraction of tin

- Tin is reduced in a similar method to iron; its ore  $\text{SnO}_2$  is heated with iron in a furnace:  $\text{SnO}_2 + 2\text{C} \rightarrow \text{Sn} + 2\text{CO}$

### (v) Extraction of gold

- Gold exists naturally as an element; its extraction consists of concentrating the gold by separating it from the surrounding rock
- Gold is unreactive but it can be converted into a soluble compound in the presence of oxygen and cyanide ions:

## Unit 9 – Metals and their Compounds

- This compound can be converted back into gold by reaction with carbon; the gold is formed on the surface of the carbon



### Test your knowledge 6.2: Extraction of Metals

- State three factors which determine how a metal is extracted from its ore
- Explain why the extraction of aluminium is expensive
- Explain why aluminium oxide is dissolved in molten cryolite
- Write the equations for the electrode reactions taking place during the electrolysis of aluminium oxide
- Explain why the anodes have to be regularly replaced
- Write two equations to show how carbon monoxide is produced in a blast furnace
- Write an equation for the reaction of tin (IV) oxide and carbon monoxide
- Write an equation for the reaction of iron (III) oxide and carbon monoxide
- Explain why limestone is added to the blast furnace
- Explain how excess carbon is removed from the iron produced in the blast furnace
- Explain why aluminium cannot be produced in the blast furnace
- State one useful by-product of the blast furnace and explain why it is useful
- Explain why iron is easier to recycle than aluminium or tin
- Explain why the extraction of gold is difficult, even though it occurs naturally
- Outline the stages in the extraction of gold (no equations required)

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

## e) Uses of Metals and Metalloids

- The uses and applications of different metals depends on their physical and chemical properties:
  - **Sodium** is too reactive to be useful in elemental form; due to its reactivity, it can be used to produce titanium from titanium chloride; this is used in the extraction of titanium from its ore:  $\text{TiCl}_4 + 4\text{Na} \rightarrow 4\text{NaCl} + \text{Ti}$
  - **Calcium** is also too reactive to be useful in elemental form; like sodium, its reactivity makes it useful in extracting other elements from their ores; calcium is used in the extraction of uranium from its ore
  - **Aluminium** is very useful in elemental form due to its strength, low density and resistance to corrosion (due to its protective oxide layer); it has a lower density than iron and doesn't corrode as easily; it is also non-toxic so can be used in food containers; aluminium is used in aeroplane parts, cans, foils and kitchen utensils; aluminium is often combined with other elements to make alloys which improve its properties (silicon is added to make it harder and stronger; magnesium is added to reduce its density)
  - **Tin** forms a protective oxide layer and is often used to coat other metals to prevent corrosion; tin combines with copper to make the alloy **bronze**; this is used to make statues and ornaments, as well as coins; tin combines with **lead** to make the alloy **solder**; this has a low melting point and is useful for welding electrical components together
  - **Iron** is very strong, especially when small amounts of carbon are added to make **steel**; steel is used to make steel girders for construction and is used in reinforced concrete; it is also used in bridges, pylons, bicycles, cars and pipes; because of its magnetic properties it is also used to make magnets
  - **Gold** is very unreactive and also soft; it is used in jewellery as it stays shiny for a long time; it is also stored as currency
  - **Copper** is unreactive and very malleable so is used to make coins and pipes; it is a good conductor of electricity so it is used to make electrical wires; when mixed with zinc it makes the alloy **brass**; brass has antibacterial properties and it used in taps and door handles, especially in bathrooms; it is shiny and resists corrosion and is often used to coat ornaments; it is also highly sonorous and so used to make musical instruments (trumpets, trumbones, saxophones and horns are known as "brass" instruments); when copper is mixed with tin it makes the alloy **bronze**; bronze is very malleable and used in coins and sculptures; brass and bronze are both unreactive, like copper
- **Metalloids** (eg silicon and boron) are semi-conductors; they conduct poorly at low temperatures but better at high temperatures; for this reason they are very useful in making electronic components; they are also added to metals to make them harder (less metallic)



### Test your knowledge 7.1: Uses of metals and metalloids

- (a) Give the name of the alloys formed by mixing the following elements:  
 (i) copper and tin; (ii) copper and zinc; (iii) tin and lead; (iv) iron and carbon
- (b) Complete the following table to show the at least one common use of the following metals and metalloids; in each case, identify the property which makes the material suitable for its use

Material	Use	Property
aluminium		
tin		
solder		
steel		
gold		
copper		
brass		
silicon		

## f) Properties and Reactions of Metal Compounds

### (i) Compounds of sodium

- All compounds of sodium are soluble in water; like most metals, it forms basic oxides and hydroxides and it can also form salts
- Sodium ions ( $\text{Na}^+$ ) are colourless; they form colourless solutions and white solids, unless the anion in the compound is coloured; this is true of all s and p-block metal ions
- The most common salts of sodium are:
  - **sodium chloride ( $\text{NaCl}$ )** is a salt of a strong acid and a strong base; it is therefore neutral; it is commonly known as salt, common salt or table salt; it is added as flavouring to food and as a source of  $\text{Na}^+$  ions; it is also placed on roads in cold weather to stop water freezing (all dissolved solutes lower the freezing point of water) as it is very cheap and non-toxic
  - **sodium nitrate ( $\text{NaNO}_3$ )** is a salt of a strong acid and a strong base; it is therefore neutral; it is commonly known as Chile saltpetre; it is an important source of nitrate ions which are important in fertilisers
  - **sodium sulphate ( $\text{Na}_2\text{SO}_4$ )** is a salt of a strong acid and a strong base; it is therefore neutral; it is mainly used in detergents
- **sodium hydroxide ( $\text{NaOH}$ )** is a strong base due to the  $\text{OH}^-$  ion; it is also known as caustic soda; it is mainly used to neutralise fatty acids to make soaps; it is also used in the manufacture of paper

### (ii) Compounds of calcium

- Like most metals, calcium forms basic oxides and hydroxides and it can also form salts
- Calcium ions ( $\text{Ca}^{2+}$ ) are colourless; they form colourless solutions and white solids, unless the anion in the compound is coloured; this is true of all s and p-block metal ions
- The most common salts of calcium are:
  - **Calcium chloride ( $\text{CaCl}_2$ )** is a salt of a strong acid and a strong base; it is therefore neutral; it is highly soluble in water even at low temperatures and is very effective at reducing the freezing point of water, hence preventing freezing; it is a very useful de-icer; it is also hygroscopic and so is used as a drying agent and is also placed on dusty roads so the road holds water better and gets less dusty, which preserves the road; it is also a water softener because the presence of  $\text{Ca}^{2+}$  ions stops  $\text{CO}_3^{2-}$  and  $\text{SO}_4^{2-}$  ions dissolving in water
  - **Calcium sulphate ( $\text{CaSO}_4$ )** is a salt of a strong acid and a strong base; it is therefore neutral; it is insoluble in water but is most stable in its hydrated form:  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$  – known as gypsum; the anhydrous form ( $\text{CaSO}_4$ ) and another hydrated form  $\text{CaSO}_4 \cdot 0.5\text{H}_2\text{O}$  are both hygroscopic, which makes them useful as drying agents;  $\text{CaSO}_4 \cdot 0.5\text{H}_2\text{O}$  is known as plaster of Paris – it forms a mouldable paste with water which hardens into  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$  when it dries; this is the most useful application of  $\text{CaSO}_4$
- **Calcium oxide ( $\text{CaO}$ )** is a base due to the  $\text{O}^{2-}$  ion; the  $\text{O}^{2-}$  ion is a strong base but  $\text{CaO}$  is only slightly soluble in water so is considered a weak base;  $\text{CaO}$  absorbs water readily to form  $\text{Ca}(\text{OH})_2$ :  $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2$ ; it is therefore hygroscopic and a good drying agent;  $\text{CaO}$  is used to remove impurities in the manufacture of steel; it reacts with acidic oxides left in the iron (mainly  $\text{SiO}_2$ ) to produce salts: eg  $\text{CaO} + \text{SiO}_2 \rightarrow \text{CaSiO}_3$



## Unit 9 – Metals and their Compounds

- **Calcium hydroxide (Ca(OH)<sub>2</sub>)** is a base due to the OH<sup>-</sup> ion; it is only slightly soluble in water and a saturated solution of Ca(OH)<sub>2</sub> in water is known as limewater and is used to test for carbon dioxide; it is used in water treatment - it is added to water to reduce its acidity and the precipitate of Ca(OH)<sub>2</sub> also acts as an effective water filter; Ca(OH)<sub>2</sub> is used in the production of cement and in the production of paper; it can also be heated to convert it into CaO:  $\text{Ca(OH)}_2 \rightarrow \text{CaO} + \text{H}_2\text{O}$



### Test your knowledge 7.2: Uses of the hydroxides, oxides and salts of sodium and calcium

- (a) Name the three most common sodium salts and state a use for each one
- (b) Name the two most common calcium salts and state a use for each one
- (c) State two uses for each of sodium hydroxide, calcium oxide and calcium hydroxide

## Lesson 8 – Why are metals and their compounds useful (part II)?

### (iii) Compounds of d-block metals (Cu, Zn and Fe)

- Most compounds of copper, zinc, iron and other d-block metals are bases or salts
- **Copper chloride (CuCl<sub>2</sub>)** is a salt of a strong acid and a relatively weak base
  - it is therefore very slightly acidic due to the hydrolysis of the Cu<sup>2+</sup> ion
  - it is a green solid and dissolves in water to form a blue-green solution
  - it is used as a catalyst in a variety of industrial processes
- **Copper sulphate (CuSO<sub>4</sub>)** is a salt of a strong acid and a relatively weak base
  - it is therefore very slightly acidic due to the hydrolysis of the Cu<sup>2+</sup> ion
  - it is a blue solid in hydrated form and dissolves in water to form a blue solution
  - it is slightly toxic and is mainly used as a fungicide
  - it is also present in a number of chemical reagents, mainly Benedict's solution and Fehling's solution, which are used to test for reducing sugars in organic chemistry; Cu<sup>2+</sup> is a mild oxidising agent as it can be reduced to copper (I) oxide (Cu<sub>2</sub>O), which is an orange solid
  - it is also used in Biuret's solution as a test for amines
- **Copper oxide (CuO)** is a base due to the O<sup>2-</sup> ion
  - the O<sup>2-</sup> ion is a strong base but CuO is insoluble in water so is considered a weak base
  - it is a black solid and it is used in pigments
  - it is also used in the disposal of some toxic compounds; it is a weak oxidising agent as it can be reduced to copper metal and can oxidise some toxic chemicals into less harmful ones

## Unit 9 – Metals and their Compounds

- d-block ions have the ability to form dative covalent bonds with other molecules or anions and form **complex ions**:
  - any molecule or anion which can form a dative covalent bond with a metal ion is called a **ligand**
  - a complex ion consists of a metal ion datively bonded to one or more ligands; the ligand provides both of the electrons needed for the covalent bond
  - two examples of complex ions are  $[\text{Cu}(\text{NH}_3)_4]^{2+}$  and  $[\text{Zn}(\text{NH}_3)_6]^{2+}$ ; they are formed when a solution containing  $\text{Cu}^{2+}$  or  $\text{Zn}^{2+}$  ions reacts with excess ammonia; these reactions take place in two stages:  
ammonia solution is an alkali and contains hydroxide ions ( $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$ ); so when ammonia solution is added to solutions containing  $\text{Cu}^{2+}$  or  $\text{Zn}^{2+}$  ions, a blue or white precipitate is formed:  
$$\text{Cu}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s}) \text{ (blue)} \quad \text{Zn}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Zn}(\text{OH})_2(\text{s}) \text{ (white)}$$
if more ammonia is added, a complex ion is formed and the precipitate redissolves to form a deep blue or a colourless solution:  
$$\text{Cu}(\text{OH})_2(\text{s}) + 4\text{NH}_3(\text{aq}) \rightarrow [\text{Cu}(\text{NH}_3)_4]^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \quad \text{Zn}(\text{OH})_2(\text{s}) + 6\text{NH}_3(\text{aq}) \rightarrow [\text{Zn}(\text{NH}_3)_6]^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$$

(deep blue)  (colourless)
  - the reactions of  $\text{Cu}^{2+}$  and  $\text{Zn}^{2+}$  with aqueous ammonia can be used to confirm their presence in aqueous solution (qualitative analysis)
- The formation of complex ions is the reason why many compounds of transition metals are coloured
  - copper ions ( $\text{Cu}^{2+}$ ), iron (II) ions ( $\text{Fe}^{2+}$ ) and iron (III) ions are all coloured
  - complex ions in which the central metal ion has partially filled d-orbitals (ie 1 – 9 electrons in a d-subshell) have the ability to absorb visible light
  - most compounds containing  $\text{Cu}^{2+}$  are blue or turquoise
  - most compounds containing  $\text{Fe}^{2+}$  are green
  - most compounds containing  $\text{Fe}^{3+}$  are orange or brown
  - the exact colour depends on the anion and on whether the compound is in solid or aqueous form
  - zinc ions ( $\text{Zn}^{2+}$ ) are colourless, like the ions of s and p-block metals; this is because the 3d subshell in  $\text{Zn}^{2+}$  is full (electronic configuration =  $[\text{Ar}]3\text{d}^{10}$ ); they form colourless solutions and white solids, unless the anion in the compound is coloured; for this reason zinc is not considered a transition metal (although it is a d-block metal)
- Water can also behave as a ligand and most d-block metal ions in solution exist in the form of hydrated complex ions: eg  $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$  and  $[\text{Cu}(\text{H}_2\text{O})_4]^{2+}$ 
  - these hydrated complex ions also exist in the solid state in the form of hydrated salts: eg  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  and  $\text{FeSO}_4 \cdot 6\text{H}_2\text{O}$
  - the presence of water as a ligand allows the electrons in partially filled d-orbitals to absorb visible light and be coloured; anhydrous compounds of transition metals contain no complex ions and so are generally not coloured  
Eg  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  (hydrated copper sulphate) is blue;  $\text{CuSO}_4$  (anhydrous copper sulphate) is white
- Anhydrous copper sulphate turns from white to blue when water is added to it:  
$$\text{CuSO}_4(\text{s}) + 5\text{H}_2\text{O}(\text{l}) \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O}(\text{s})$$

white  blue

**This reaction can be used as a test for water:** add a few drops of the liquid to anhydrous copper sulphate; if water is present, the white solid should turn blue

## Unit 9 – Metals and their Compounds



### Practical 8.1: Forming complex ions by reacting metal ions with excess ammonia

You are given seven solutions: lead (II) nitrate, calcium chloride, iron (II) sulphate, iron (III) sulphate, zinc sulphate, aluminium sulphate and copper sulphate; you are also given a solution of ammonia

- Pour 2 cm<sup>3</sup> of one of the above solutions into a test tube
- Add a few drops of NH<sub>3</sub>, then add NH<sub>3</sub> until the test tube is almost full; note your observations
- Repeat steps (a) and (b) using the other solutions
- Record all of your observations in the following table:

Cation present in solution	Observation on adding a few drops of ammonia	Observation on adding excess ammonia
Pb <sup>2+</sup>		
Ca <sup>2+</sup>		
Fe <sup>2+</sup>		
Fe <sup>3+</sup>		
Zn <sup>2+</sup>		
Al <sup>3+</sup>		
Cu <sup>2+</sup>		

**Write ionic equations for the precipitation reactions taking place when a few drops of ammonia are added.**  
**Write ionic equations for any complex formation reactions taking place when excess ammonia is added.**  
**Which cations have different reactions with excess NH<sub>3</sub> and excess NaOH?**



### Practical 8.2: React anhydrous copper sulphate with water

- Using a spatula, place 2 – 3 g of anhydrous copper sulphate on an evaporating dish
- Add a few drops of water; note any change in colour
- Repeat steps (a) and (b) but add ethanol and paraffin instead of water; note any change in colour

**What type of reaction is taking place when water is added to anhydrous copper sulphate?**



### Test your knowledge 8.3: Properties and reactions of compounds of d-block metals

- Give the formula of the two most common copper salts and give one use of each
- Give the formula of the most common base of copper and state one use for it
- Define the terms “complex ion” and “ligand”; give one example of a complex ion
- Explain why complex ions of d-block metals are often coloured
- Explain why zinc compounds are colourless
- Explain why anhydrous copper sulphate is colourless
- Write an equation for the reaction between Cu<sup>2+</sup> and excess ammonia and state what you would see
- Write an equation for the reaction between Zn<sup>2+</sup> and excess ammonia and state what you would see
- Describe a chemical test for water and write an equation for the reaction taking place

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

## g) Qualitative Analysis Part 4 – Complex Formation Reactions



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

- What is qualitative analysis?
- Which cations form precipitates with aqueous sodium hydroxide? What are the colours of these precipitates?
- Which of these precipitates dissolve in excess aqueous sodium hydroxide?

- Some simple qualitative tests for cations, anions and gases were covered in Unit 5 – Acids, Bases and Salts, Unit 6 – Redox Reactions and Unit 8 – Solubility and Precipitation
- Complex formation reactions are useful in qualitative analysis;  $\text{Fe}^{2+}$ ,  $\text{Fe}^{3+}$ ,  $\text{Ca}^{2+}$ ,  $\text{Al}^{3+}$ ,  $\text{Pb}^{2+}$ ,  $\text{Cu}^{2+}$  and  $\text{Zn}^{2+}$  all give insoluble hydroxide precipitates when added to ammonia but only  $\text{Cu}(\text{OH})_2$  and  $\text{Zn}(\text{OH})_2$  dissolve in excess ammonia to give a complex ion



### Practical 9.2: Qualitative Analysis Part 4: use complex formation reactions to identify cations in solution

You are given four colourless solutions, labelled A, B, C and D; each solution contains one cation from the following:  $\text{Ca}^{2+}$ ,  $\text{Zn}^{2+}$ ,  $\text{Pb}^{2+}$  and  $\text{Al}^{3+}$

- (a) Pour 2 cm<sup>3</sup> of solution A into a test tube
- (b) Add a few drops of NaOH, then adding NaOH until the test tube is almost full; note your observations
- (c) Repeat steps (a) to (b) using solutions B – D
- (d) Pour 2 cm<sup>3</sup> of solution A into a test tube
- (e) Add a few drops of NH<sub>3</sub>, then adding NH<sub>3</sub> until the test tube is almost full; note your observations
- (f) Repeat steps (a) to (b) using solutions B – D

Use your observations to identify which cation is in which solution; present your answers in a table similar to the one below:

Solution	Observations				Cation Present
	Few drops NaOH	Excess NaOH	Few drops NH <sub>3</sub>	Excess NH <sub>3</sub>	
A					
B					
C					
D					

Which two cations cannot be distinguished by this combination of tests? How would you distinguish between solutions containing these cations?

## Unit 9 – Metals and their Compounds



### Test your knowledge 9.3: Qualitative Analysis Part 4

- (a) Describe a chemical test to distinguish between solutions of aluminium sulphate and zinc sulphate

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



### 10.1 END-OF-UNIT QUIZ UNIT 9 – METALS AND THEIR COMPOUNDS

1. Describe the bonding in metals and hence explain why metals are malleable and ductile.
2. State two other physical properties of metals and explain why metals have these properties.
3. Write equations for the following reactions and state the type of reaction occurring:
  - (a) Magnesium with dilute nitric acid
  - (b) Sodium with water
  - (c) Zinc with copper sulphate
4. Explain why zinc reacts with copper sulphate but copper does not react with zinc sulphate.
5. Describe how aluminium is extracted from its ore.
6. State why brass and solder are useful alloys and state which pure metals they are made from.
7. Explain why aqueous solutions of zinc salts are colourless but aqueous solutions of copper salts are coloured.
8. Explain the difference between a d-block metal and a transition metal.
9. Explain why transition metals can form variable oxidation states.
10. State what you would see when dilute ammonia solution is added gradually to a solution containing  $\text{Zn}^{2+}$  ions. Write equations for any reactions occurring.
11. Explain how you would distinguish between aqueous solutions of zinc nitrate and aluminium nitrate.
12. Describe a chemical test for water.