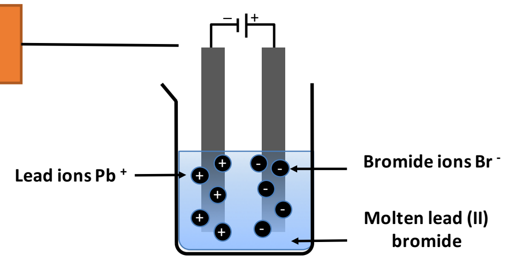
Chemistry SC10-13 Electrolysis

Electrolytic processes

Process of electrolysis	*Splitting up using electricity*	When an ionic compound is melted or dissolved in water, the ions are free to move. These are then able to conduct electricity and are called electrolytes. Passing an electric current though electrolytes causes the ions to move to the electrodes.
Electrode	*Anode* *Cathode*	The positive electrode is called the anode. The negative electrode is called the cathode.
Where do the ions go?	*Cations* *Anions*	Cations are positive ions and they move to the negative cathode. Anions are negative ions and they move to the positive anode.

Oxidation Is Loss, Reduction Is Gain

Extracting metals using electrolysis
Metals can be extracted from molten compounds using electrolysis.
This process is used when the metal is too reactive to be extracted by reduction with carbon.
The process is expensive due to large amounts of energy needed to produce the electrical current. *Example: aluminium is extracted in this way.*



Higher tier: **You can display what is happening at each electrode using half-equations:**

At the cathode: $Pb^{2+} + 2e^- \rightarrow Pb$

At the anode: $2Br^- \rightarrow Br_2 + 2e^-$

Electrolysis of aqueous solutions

The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved.

At the negative electrode	Metal will be produced on the electrode if it is less reactive than hydrogen. Hydrogen will be produced if the metal is more reactive than hydrogen.
At the positive electrode	Oxygen is formed at positive electrode. If you have a halide ion (Cl^-, I^-, Br^-) then you will get chlorine, bromine or iodine formed at that electrode.

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Oxidation **I**s **L**oss (of electrons) **R**eduction **I**s **G**ain (of electrons)

Chemistry SC 10-13

Obtaining and using metals

Oxidation and reduction in terms of electrons (HT ONLY)

Ionic half equations (HT only)		
For displacement reactions	*Ionic half equations show what happens to each of the reactants during reactions*	For example: The ionic equation for the reaction between iron and copper (II) ions is: **$Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$** The half-equation for iron (II) is: **$Fe \rightarrow Fe^{2+} + 2e^-$** The half-equation for copper (II) ions is: **$Cu^{2+} + 2e^- \rightarrow Cu$**

Reactions of acids and metals

Acids react with some metals to produce salts and hydrogen.

Reactions with acids	*metal + acid → metal salt + hydrogen*	**magnesium + hydrochloric acid → magnesium chloride + hydrogen** **zinc + sulfuric acid → zinc sulfate + hydrogen**

HT ONLY: Reactions between metals and acids are redox reactions as the metal donates electrons to the hydrogen ions. This displaces hydrogen as a gas while the metal ions are left in the solution.

Extraction of metals and reduction

Unreactive metals, such as gold, are found in the Earth as the metal itself. They can be mined from the ground. More reactive metals are obtained by displacement or electrolysis.

Extraction using carbon	
Metals less reactive than carbon can be extracted from their oxides by reduction.	For example: **zinc oxide + carbon → zinc + carbon dioxide**

Obtaining and using metals

Metals ores	*These resources are limited*	Copper ores especially are becoming sparse. New ways of extracting copper from low-grade ores are being developed.
Phytomining	*Plants absorb metal compounds*	These plants are then harvested and burned; their ash contains the metal compounds.
Bioleaching	*Bacteria is used to produce leachate solutions that contain metal compounds*	The metal compounds can be processed to obtain the metal from it e.g. copper can be obtained from its compounds by displacement or electrolysis.

Metal oxides

Metals and oxygen	*Metals react with oxygen to form metal oxides*	**magnesium + oxygen → magnesium oxide** *$2Mg + O_2 \rightarrow 2MgO$*
Reduction	*This is when oxygen is removed from a compound during a reaction*	e.g. metal oxides reacting with hydrogen, extracting low reactivity metals
Oxidation	*This is when oxygen is gained by a compound during a reaction*	e.g. metals reacting with oxygen, rusting of iron

The reactivity series

	Reactions with water	Reactions with acid
Group 1 metals	*Reactions get more vigorous as you go down the group*	*Reactions get more vigorous as you go down the group*
Group 2 metals	*Do not react with water*	*Observable reactions include fizzing and temperature increases*
Zinc, iron and copper	*Do not react with water*	*Zinc and iron react slowly with acid. Copper does not react with acid.*

Metals form positive ions when they react	*The reactivity of a metal is related to its tendency to form positive ions*	The reactivity series arranges metals in order of their reactivity (their tendency to form positive ions).
Carbon and hydrogen	*<u>Carbon</u> and <u>hydrogen</u> are non-metals but are included in the reactivity series*	These two non-metals are included in the reactivity series as they can be used to extract some metals from their ores, depending on their reactivity.
Displacement	*A more reactive metal can displace a less reactive metal from a compound.*	Silver nitrate + Sodium chloride → Sodium nitrate + Silver chloride

potassium	most reactive	K
sodium		Na
calcium		Ca
magnesium		Mg
aluminium		Al
carbon		C
zinc		Zn
iron		Fe
tin		Sn
lead		Pb
hydrogen		H
copper		Cu
silver		Ag
gold		Au
platinum	least reactive	Pt

Chemistry SC 10-13

Reversible reactions and equilibria

Energy changes and reversible reactions

If one direction of a reversible reaction is exothermic, the opposite direction is endothermic. The same amount of energy is transferred in each case.

For example: Hydrated copper sulfate $\underset{\text{exothe}}{\overset{\text{endothermic}}{\rightleftharpoons}}$ Anhydrous copper + Water sulfate

Reversible reactions

Reversible reactions	In some chemical reactions, the products can react again to re-form the reactants.
Representing reversible reactions	$A + B \rightleftharpoons C + D$
The direction	The direction of reversible reactions can be changed by changing conditions: $A + B \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} C + D$

Equilibria

Equilibrium in reversible reactions	When a reversible reaction occurs in apparatus which prevents the escape of reactants and products, equilibrium is reached when the forward and reverse reactions occur exactly at the same rate.

Changing conditions and equilibrium (HT)

The relative amounts of reactants and products at equilibrium depend on the conditions of the reaction.

Le Chatelier's Principles	States that when a system experiences a disturbance (change in condition), it will respond to restore a new equilibrium state.
Changing concentration	If the concentration of a reactant is increased, more products will be formed . If the concentration of a product is decreased, more reactants will react.
Changing temperature	If the temperature of a system at equilibrium is increased: - Exothermic reaction = products decrease - Endothermic reaction = products increase
Changing pressure (gaseous reactions)	For a gaseous system at equilibrium: - Pressure increase = equilibrium position shifts to side of equation with smaller number of molecules. - Pressure decrease = equilibrium position shifts to side of equation with larger number of molecules.

The Haber process

The Haber process	This process uses nitrogen from the air and hydrogen from natural gas to form ammonia. The reaction is reversible and uses optimum conditions and a catalyst in order to reach dynamic equilibrium.
Optimum temperature	The optimum temperature for the Haber process is 450^{o}C.
Optimum pressure	The optimum pressure for the Haber process is 200 atmospheres.
The use of a catalyst	The Haber process uses an iron catalyst. This does not alter the position of the equilibrium but it does increase the rate of the reaction.

Life cycle assessment and recycling

Life cycle assessment

LCAS	*Life cycle assessments are carried out to assess the environmental impact of products*	They are assessed at these stages: - Extraction and processing raw materials - Manufacturing and packaging - Use and operation during lifetime - Disposal
Values	*Allocating numerical values to pollutant effects is difficult*	Value judgments are allocated to the effects of pollutants so LCA is not a purely objective process.

Ways of reducing the use of resources

Reduce, reuse and recycle	*This strategy reduces the use of limited resources*	This, therefore, reduces energy sources being used, reduces waste (landfill) and reduces environmental impacts.
Limited raw materials	*Used for metals, glass, building materials, plastics and clay ceramics*	Most of the energy required for these processes comes from limited resources. Obtaining raw materials from the Earth by quarrying and mining causes environmental impacts.
Reusing and recycling	*Metals can be recycled by melting and recasting/reforming*	Glass bottles can be reused. They are crushed and melted to make different glass products. Products that cannot be reused are recycled.

Chemistry SC10-13

Dynamic equilibria

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Optimum conditions	The optimum temperature for the Haber process is 450^{o}C and optimum pressure is 200 atmospheres. These are economically viable conditions as they produce the best yield to cost ratio.

Fertilisers

NPK fertilisers	*These contain nitrogen, phosphorous and potassium*	Formulations of various salts containing appropriate percentages of the elements.
Fertiliser examples	*Potassium chloride, potassium sulfate and phosphate rock are obtained by mining*	Phosphate rock needs to be treated with an acid to produce a soluble salt which is then used as a fertiliser. Ammonia can be used to manufacture ammonium salts and nitric acid.

Transition metals, alloys and corrosion

Properties and uses of metals

Metal	*Properties*	Uses
Aluminium	*Low density (lightweight), layer of oxides at surface (corrosion resistant)*	Aluminium cans, cooking foil, saucepans.
Copper	*Good electrical and thermal conductor, flexible*	Saucepans, electrical wiring.
Gold	*Unreactive*	Jewellery, coins.

Magnalium (Aluminium and magnesium alloy)	Aircraft and car parts.
Brass (copper and zinc alloy)	Used in electrical fittings.

Oxidation

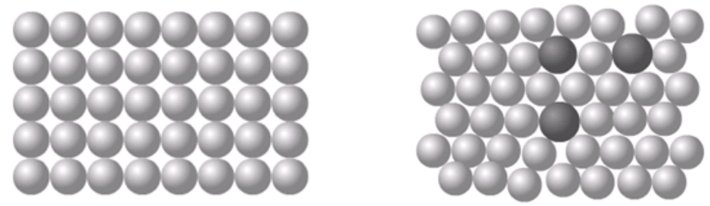
Corrosion	*The destruction of materials by chemical reactions with substances in the environment*	An example of this is iron rusting; iron reacts with oxygen from the air to form iron oxide (rust) water needs to be present for iron to rust.
Preventing corrosion	*Coatings can be added to metals to act as a barrier*	Examples of this are greasing, painting and electroplating. Aluminium has an oxide coating that protects the metal from further corrosion.
Sacrificial corrosion	*When a more reactive metal is used to coat a less reactive metal*	This means that the coating will react with the air and not the underlying metal. An example of this is zinc used to galvanise iron.
Electroplating	*Used to improve the appearance and/or resistance to corrosion*	Electrolysis is used to reduce metal cations so they form a thin layer at the cathode.

Properties of metals and alloys

High melting and boiling points	This is due to the strong metallic bonds.
Pure metals can be bent and shaped	Atoms are arranged in layers that can slide over each other.

Alloys	*Mixture of two or more elements at least one of which is a metal*	Harder than pure metals because atoms of different sizes disrupt the layers so they cannot slide over each other.

Pure metal / Alloy



Transition metals

Transition metals	*Most metals are transition metals*	• High melting points • High density • They form coloured compounds • They can be used as catalysts (without being used up)