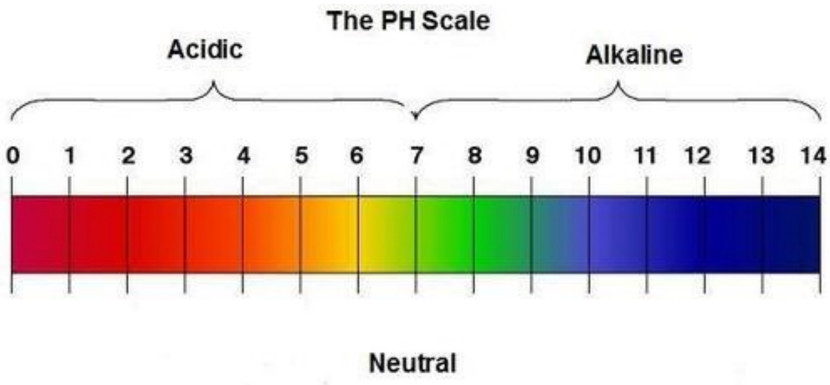
Acids & Bases:

pH scale:

Low pH = acidic

High pH = alkaline

Neutral = pH 7



Measuring pH:

Wide range indicators: mixture of dyes that gradually change colour over a range of pHs -> useful for estimating e.g. Universal indicator

pH probes (in a solution): attach to a pH meter to measure pH electronically on a digital display (numerical value = more accurate)

Neutralisation:

Acids = form aqueous solutions with pH <7, forming H+ ions in H2O

Base = pH >7

Alkali = base that dissolves in H20 to form a solution with pH >7, form OH+ ions in H2O

Acid + base -> salt + water

H+ + OH- -> H20

Gives a neutral (7) pH

Titrations: calculate exact contraction of an acid/alkali needed to form a neutralisation reaction

Single indicators needed:

Show single colour changes e.g. using phenolphthalein or litmus NOT universal indicators as they cover wide pH ranges (as it's a mixture of indicators) to get the end-point: a sudden colour change

Strong acids & Weak acids:

Acids:
Ionise in aqueous solutions, producing H^+ ions

Strong acids: ionise (dissociate) COMPLETELY in water, releasing H^+ ions

Weak acids: ionise (dissociate) PARTIALLY in water, to release H^+ ions; it's a reversible reaction which sets up an equilibrium between the undissociated & dissociated acid (only a few H^+ ions are released = favours the reactant)

Higher concentration of H^+ ions = faster rate of reaction so strong acids are more reactive than weak acids of the same concentration

pH is a measure of the concentration of H^+ ions in the solution:

Factor H^+ ion concentration changes by = 10^{-x}

For every decrease of 1 on the pH scale, the concentration of H^+ ions increases by a factor of 10
So...
For every decrease of 2 on the pH scale, the concentration of H^+ ions increases by a factor of 100 ECT
Thus the pH of a strong acid is less than of a weak acid if they have the same concentration

Acid strength = how much the H^+ ions ionise in water
Concentration = how much acid in a given volume of water
More acid in a fixed volume of water = more concentrated
So can have a concentrated but weak acid OR a dilute but strong acid!
pH decreases with increasing acid concentration (dissociation of H^+ ions) regardless of acid strength

Reactions of acids with bases:

Acid + metal oxide -> salt + water

Acid + metal hydroxide -> salt + water

Even bases that dont dissolve in water react in neutralisation reactions, shown above

Acid + metal carbonate -> salt + water + carbon dioxide

Hydrochloric acid: chloride
Sulphuric acid: sulphate
Nitric acid: nitrate

E.g. hydrochloric acid + sodium carbonate -> sodium chloride + water + carbon dioxide

Reactivity series:

Metals: lose electrons so form positive ions
Non-metals: gain electrons so form negative ions
= how easily they do so determines their reactivity

Metals react with water & acids

Acid + metal -> salt + hydrogen
Speed of reaction determined by rate of hydrogen bubbles given off

Measuring temperature change of the same mass of reactant & surface area = more reactive metals have a greater temp change

Metal + water -> metal hydroxide + hydrogen

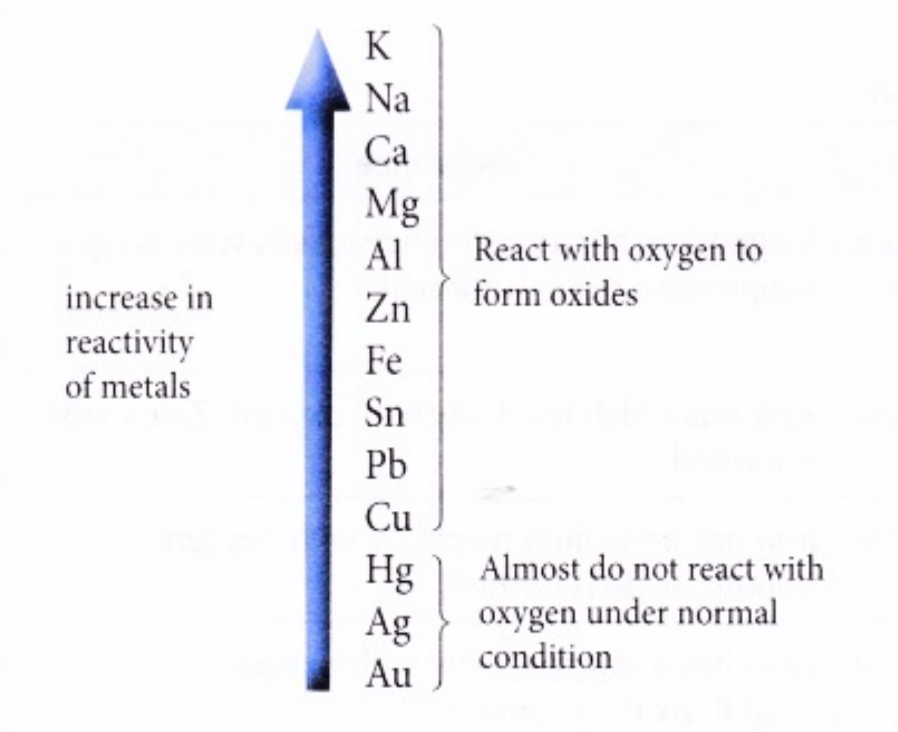


Figure *The reactivity series of metals towards oxygen*

Separating metals from metal oxides:

Iron & oxygen react = iron oxide via OXIDATION in ores
So use REDUCTION to separate a metal from its oxide

OILRIG= Oxidation is gain, Reduction is loss OF ELECTRONS

Extracting metals from their ores:

Reduction with carbon: oxygen is removed from the ore to react with carbon (its oxidised) whilst the metal is reduced

Metals higher than carbon on the reactivity series are extracted via ELECTROLYSIS
Metals lower than carbon are extracted via REDUCTION
-> can only take oxygen away from metals that are less reactive in this displacement reaction

Gold = unreactive so mined in its raw form

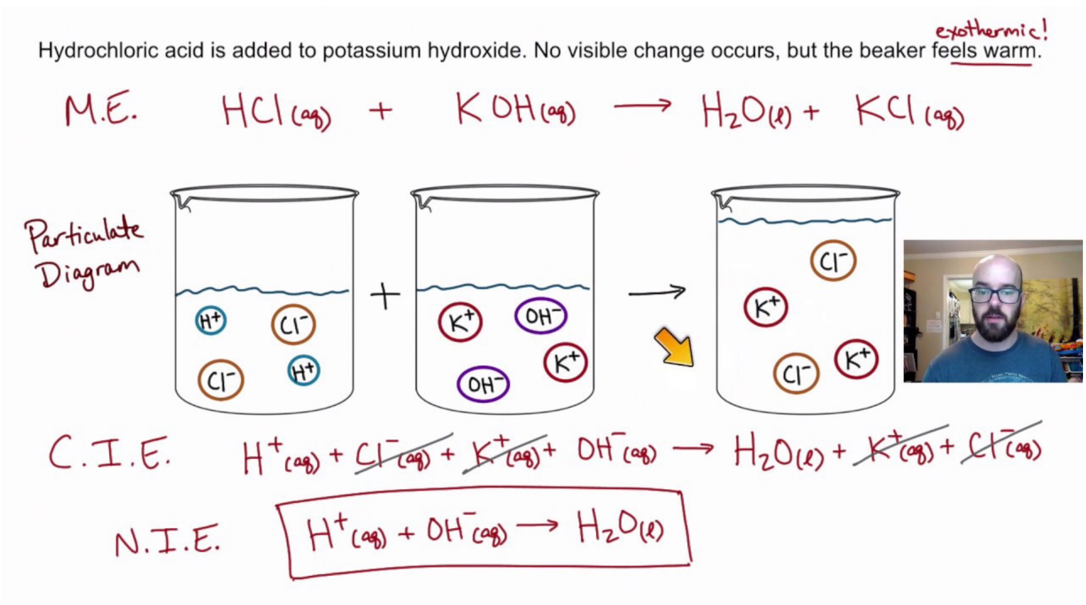
Redox reactions:

OILRIG = OXIDATION is loss, Reduction is gain (of electrons not of oxygen)

So REDOX reaction = REDuction & OXidation happen at the SAME time

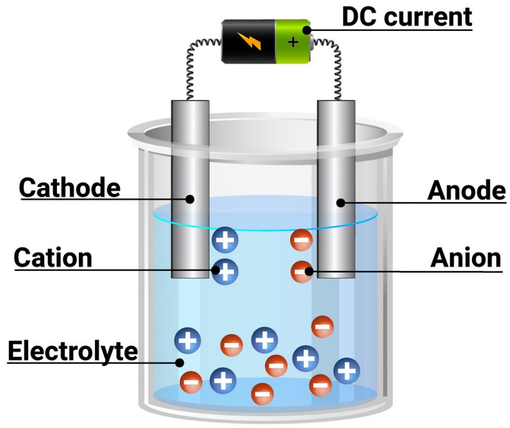
Displacement reactions: a more reactive element displaces a less reactive metal from its compound/ore
Metal ion gains electrons = is reduced
Metal atom looses electrons = is oxidised

Ionic equations: focus on the substances that are oxidised or reduced, spectator ions can be crossed out



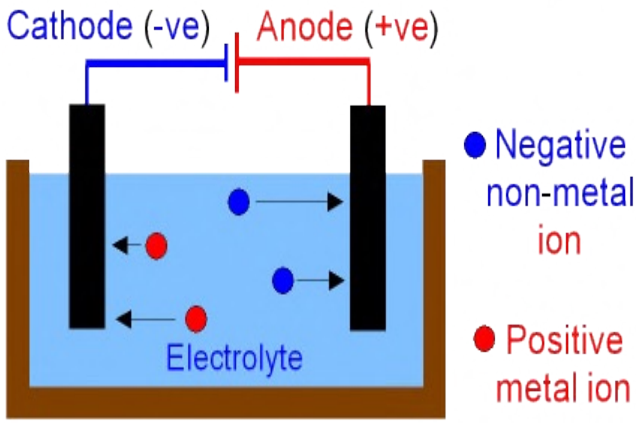
Electrolysis:

The decomposition of a compound when a current is passed through an electrolyte (molten/dissolved ionic compound)



Positive ions (cations) move to the negative electrode (cathode) & gain electrons = reduction
Negative ions (anions) move to the positive electrode (anode) & lose electrons = oxidisation

When ions lose/gain electrons, the uncharged element discharges from the electrolyte



Molten ionic compounds are electrolysed as ions move freely = conducting electricity
Positive metal ions are reduced at the cathode
Negative non-metal ions are oxidised at the anode

If a metal is more reactive than carbon = can't use displacement so use electrolysis (costly as energy intensive, in melting the ore & producing a current)

Electrolysis of aluminium oxide:

Extracted from bauxite ore
Cryolite lowers the melting point so ions can move to the electrodes (only carry current when molten)

Cations (Al^+3) gain electrons at the cathode =aluminium atoms sink

Anions (O^2-) lose electrons at the anode= oxygen atoms combine to form O2

Anode is replaced as it's made of graphite (carbon) so reacts with the oxygen created, forming CO2 gas

Extraction of aluminium using electrolysis – half reactions

Anode (+ve electrode)
$2O^{2-}_{(l)} \rightarrow O_{2\,(g)} + 4e^-$

Cathode (-ve electrode)
$Al^{3+}_{(l)} + 3e^- \rightarrow Al_{(l)}$

Overall
$2Al_2O_{3\,(l)} \rightarrow 4Al_{(l)} + 3O_{2\,(g)}$

The anode reacts to form carbon dioxide
$C + O_2 \rightarrow CO_2$

Electrolysis of aqueous solutions:

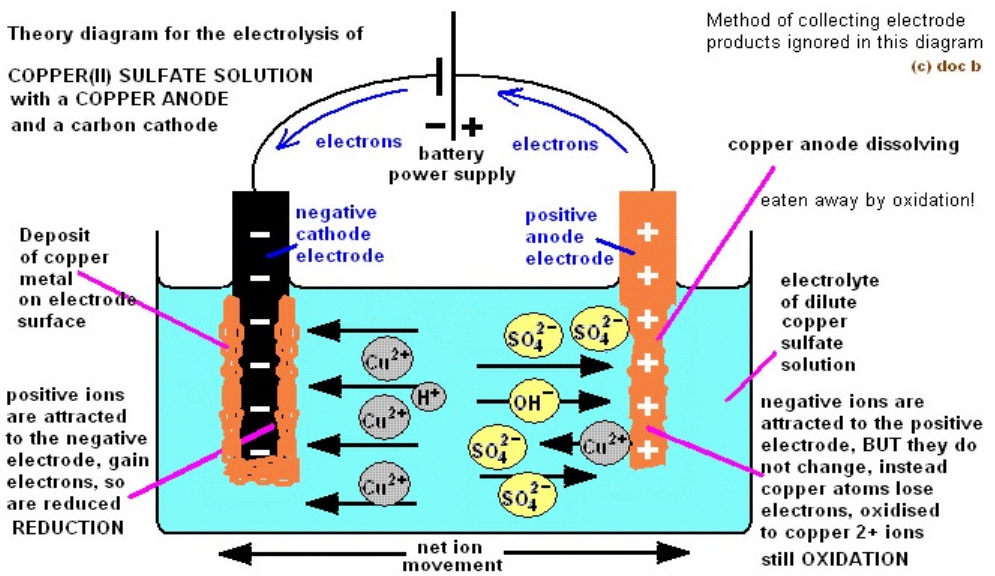
Will be OH- & H+ ions from the H2O

Cathode: hydrogen forms if element produced is more reactive than it, if less reactive a pure metal forms

Anode: if halogens form, OH- is formed, if not OH- ions are discharged & O2 is formed

E.g. in CuSO^4:
Cathode: copper is formed as is less reactive than hydrogen
Anode: oxygen formed as no halides present

Apparatus for electrolysis:



Testing which gases were produced:

Chlorine bleached damp litmus paper
Hydrogen makes a squeaky pop with a splint
Oxygen relights a glowing splint