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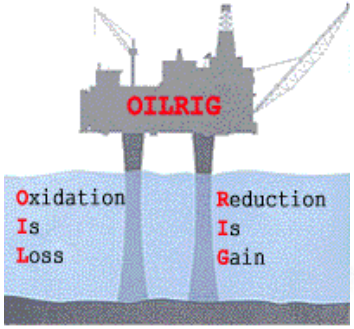
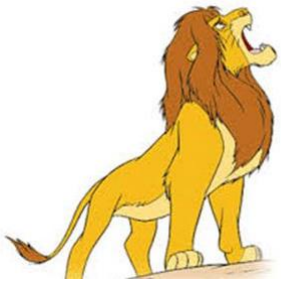




LEO says GER

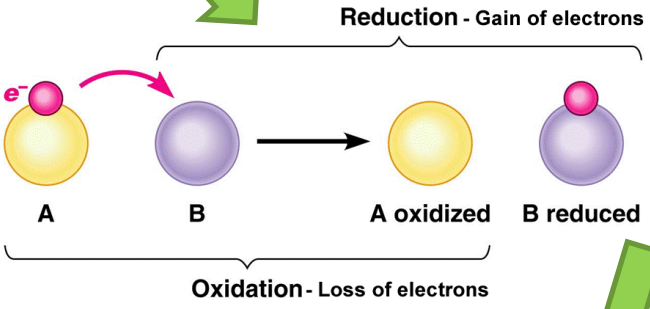
Loss Electrons = Oxidation

Gain Electrons = Reduction

Name of acid	Formula	Formula of ion	Name of salt
Hydrochloric acid	HCl	Cl^-	Chloride
Sulphuric acid	H_2SO_4	SO_4^{-2}	Sulphate
Nitric acid	HNO_3	NO_3-	Nitrate
Phosphoric acid	H_3PO_4	PO_4^{-3}	Phosphate

Key words:
Oxidation: loss of electrons
Reduction: gain of electrons
Displacement reaction: When a more reactive metal swaps places with a less reactive metal

Metal + water -> metal hydroxide + hydrogen
Metal + acid -> metal salt + hydrogen



Increasing reactivity

- Potassium
- Sodium
- Lithium
- Calcium
- Magnesium
- Aluminium
- Carbon
- Zinc
- Iron
- Hydrogen
- Copper
- Silver
- Gold

Order of reactivity	Reaction with water	Reaction with dilute acid
potassium	fizz, giving off hydrogen gas, leaving an alkaline solution of metal hydroxide	explode
sodium		
lithium		
calcium		fizz, giving off hydrogen gas and forming a salt
magnesium	very slow reaction	
aluminium		
zinc		
iron		
tin	slight reaction with steam	react slowly with warm acid
lead		
copper	no reaction, even with steam	no reaction
silver		
gold		

What affect how reactive something is?
Metals: The less electrons to lose and the further away from the nucleus as the attractive force from the nucleus is less, electrons are easier to lose so the metal is more reactive.
Non-metals: The less electrons to gain and the closer to the nucleus means the attractive force is greater.

- Metals become oxidised
- Non-metals such as oxygen become reduced when they form an ionic bond
- In this case it has also formed an oxide because it has reacted with oxygen
- Metals react easily with oxygen so often are found in oxide compounds
- If we want to remove the oxygen we need to find a more reactive metal to displace the metal

Sometimes we cant extract a metal using carbon as carbon forms another compound. When you react carbon with tungsten oxide it forms tungsten carbide. So instead we react it with hydrogen in a process called electrolysis which requires electricity.

When we extract a metal using carbon from its ore we call this reduction of oxide by carbon. The metal oxide is heated up with carbon and if carbon is higher in the reactivity series displacement happens and carbon dioxide if formed in the place of the metal oxide.

Metals are found in ores in rocks where they are in a compound with oxygen. Silver, gold and platinum are never found in an ore, as they are very unreactive.

Ionic equation:

This only shows the atoms and ions that change (are oxidised or reduced)
It does not show the ions that stay the same (the salt)

Half equation:

This shows the movement of electrons
There are 2 equations
Each equation shows the movement of electrons, which atoms and ions are oxidised and reduced

1. Reactivity Series of Metals

Potassium	most reactive
Sodium	
Calcium	
Magnesium	
Aluminium	
Carbon	
Zinc	
Iron	
Tin	
Lead	
Hydrogen	
Copper	
Silver	
Gold	
Platinum	least reactive

(elements in italics, though non-metals, have been included for comparison)

2. Formulae of Some Common Ions

Positive ions		Negative ions	
Name	Formula	Name	Formula
Hydrogen	H^+	Chloride	Cl^-
Sodium	Na^+	Bromide	Br^-
Silver	Ag^+	Fluoride	F^-
Potassium	K^+	Iodide	I^-
Lithium	Li^+	Hydroxide	OH^-
Ammonium	NH_4^+	Nitrate	NO_3^-
Barium	Ba^{2+}	Oxide	O^{2-}
Calcium	Ca^{2+}	Sulfide	S^{2-}
Copper(II)	Cu^{2+}	Sulfate	SO_4^{2-}
Magnesium	Mg^{2+}	Carbonate	CO_3^{2-}
Zinc	Zn^{2+}		
Lead	Pb^{2+}		
Iron(II)	Fe^{2+}		
Iron(III)	Fe^{3+}		
Aluminium	Al^{3+}		

$Fe_{(s)} + CuSO_{4\,(aq)} \rightarrow FeSO_4 + Cu_{(s)}$

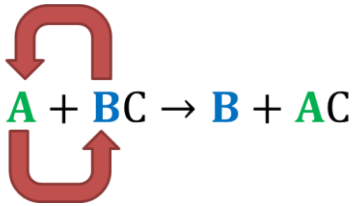
Ionic Equation

$Fe_{(s)} + Cu^{2+}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$

Half equation

$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$



Displacement reaction:

Magnesium + copper(II) sulphate -> magnesium sulphate + copper

$Mg_{(s)} + CuSO_{4\,(aq)} \rightarrow MgSO_4 + Cu_{(s)}$

- In a **ionic equation** it **only** shows the atoms that have changed in a reaction (so you just get rid of the salt part)
- A half equation is where you write 2 equations one for the oxidation reaction (metal) and one for the reduction (non-metal)

Ionic Equation

$aMg_{(s)} + Cu^{2+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + Cu_{(s)}$

Half equation

$Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e^-$

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

- When reacting an acid and an alkali you need to be able to tell that the acid has fully reacted as it is not obvious when this has happened so you need an acid base indicator and you would carry out a titration
- Carry out the titration with the indicator added to see how much acid reacts completely with the alkali
- Run that volume of acid into the solution of the alkali again, but this time without the indicator
- Then crystallise and dry the crystals slowly

pH is simply a measurement of the concentration of H^+ ions in a solution. At pH7 the concentration of H+ and OH- is equal so the solution is 'neutral.'

Acid + oxide (base) → salt + water
Acid + hydroxide (alkali) → salt + water
Acid + carbonate → salt + water + carbon dioxide

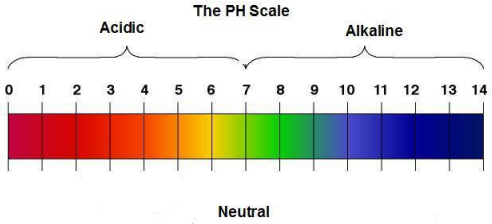
When an acid reacts with an alkali it donates H+ ions, the alkali OH- ions and H_2O is formed

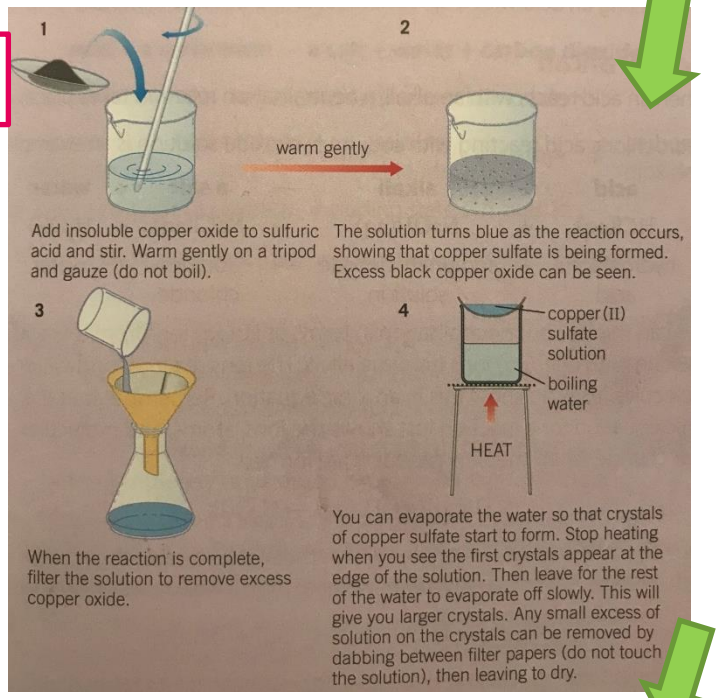
Key words:

Acids: pH of less than 7, form H+ (hydrogen) ions when in solution (dissolved in water - aq).

Base: solids that react with an acid. E.g. metals, metal oxides, metal carbonates

Alkali: a base that dissolves in water (soluble base - aq), form OH- (hydroxide) ions in solution. pH of greater than 7.





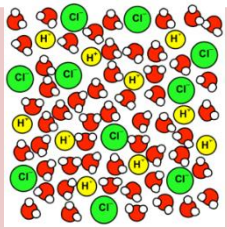
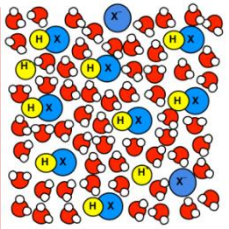
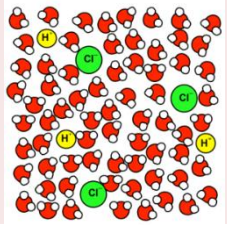
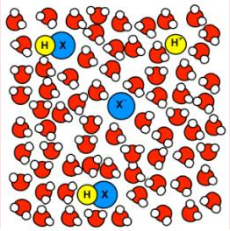
Acid + alkali = neutral
You can measure pH using universal indicator of a pH probe

The salt copper sulfate can be made by reacting copper carbonate with dilute sulfuric acid.

$CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$

Making copper sulphate

- pour a suitable volume of sulphuric acid into a suitable container
- add a small amount of copper carbonate to the acid and stir until the effervescence stops
- continue to add small amounts of copper carbonate to the acid and each time stir until any effervescence stops
- eventually when there is no reaction / effervescence when the copper carbonate is added filter the mixture to remove the excess copper carbonate
- pour the filtrate (copper nitrate solution) into an evaporating basin and heat to evaporate a small amount of the water
- leave the copper sulphate solution to crystallise
- remove the crystals from the solution remaining and dry the crystals

	STRONG (all of the molecules break down into ions)	WEAK (small fraction of the molecules break down into ions)
CONCENTRATED (a lot of acid is dissolved)		
DILUTE (small amount of acid dissolved)		

Strong acid: all the molecules are broken down into ions

Weak acid: small fraction of the molecules are broken down into ions

Dilute acid: small amount of acid dissolved

Concentrated acid: A lot of acid is dissolved

- Acids must dissolve in water to be acidic
- The more H^+ that forms in the solution the stronger the acid
- Strong acids ionise completely in solution and it is not reversible
- $HCl(aq) \rightarrow Cl^-(aq) + H^+(aq)$
- Weak acids most of the molecules stay as they are and only a relatively small proportion of the acid molecules will ionise the solution, their reaction is reversible unlike a strong acid.
- So as the molecules of a weak acid split up to for H^+ ions and negative ions, the ions recombine to form the original molecules again.
- A position of equilibrium is reached in which both whole molecules (the majority) and their ions (the minority) are present. So in ethanoic acid the following equation happens:
- $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

Concentration of $H^+(aq)$ ions in mol/dm^3		pH value
0.10	0.1×10^{-1}	1.0
0.01	0.1×10^{-2}	2.0
0.0010	0.1×10^{-3}	3.0
0.00010	0.1×10^{-4}	4.0

When you put it into standard form the number after the minus tells you the pH (just ignore the minus sign)

A concertation of 0.1 has a pH of x10 more

As the concertation decreases (H+ ions decrease) by a factor of 10 the pH goes up by 1 unit

As the concertation increase (H+ ions decrease) by a factor of 10 the pH goes down by 1 unit

1. When potassium reacts with oxygen it forms potassium oxide. What type of bond forms?
2. Potassium is in group 1 and oxygen is in group 6:
   a. How many electrons does potassium need to lose to get a full outer shell?
   b. How many electrons does oxygen need to gain to get a full outer shell?
   c. How many potassium atoms will bind to 1 oxygen atom
3. Write a balanced equation for the reaction:
4. Oxidation is where electrons are lost, which atom has been oxidised, potassium or oxygen?
5. Reduction is where atom gains electrons, which atom has be reduced, potassium or oxygen?
6. Explain the 2 factors that affect the reactivity of a metal and justify why the metal you have chosen for your answer to question 3 is the most reactive.

Complete the word equations and complete a balanced symbol equation for the following displacement reactions

1. Magnesium oxide + calcium -> calcium oxide + magnesium
2. Silver oxide + magnesium -> magnesium oxide + silver
3. Lead oxide + lithium -> lithium oxide + lead
4. Calcium oxide + magnesium -> Calcium oxide + magnesium
5. Potassium oxide + calcium -> Potassium oxide + calcium
6. Copper oxide + sodium -> sodium oxide + copper
7. Metals are often found in an ore e.g. copper oxide, suggest using ideas about the reactivity series how scientists could extract copper from its ore:
8. Recall the different ways you would know a reaction had taken place
9. Aluminium reacts with oxygen to make aluminium oxide which forms a protective layer on top of the aluminium. Explain why this means aluminium can be used outside
10. Explain why silver and gold are used to make jewellery
11. Identify and justify which metal would be best for water pipes in your home
12. Explain why potassium, lithium and sodium are stored in oil

1. What is a metal ore?
2. Explain the factors that are considered before a metal is extracted from its ore
3. Balance the reaction
4. In the **reactants** has Pb been reduced or oxidised?
5. In the **reactants** has O been reduced or oxidised?
6. In the **products** has Pb been reduced or oxidised?
7. In the **products** has C been reduced or oxidised?
8. Explain using ideas about the reactivity series why this reaction has occurred
9. Write a word equation for the extraction of copper from its ore
10. Write a balanced equation, copper in an aqueous solution has a charge of +2
11. After the reaction has copper been reduced or oxidised?
12. After the reaction has carbon been reduced of oxidised?
13. Explain how copper can be extracted from its ore:

Write an ionic and half equation for the following reactions with metals and acids

1. $Cu(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2(g)$
2. $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
3. Sodium carbonate + Hydrochloric acid.
4. Potassium hydroxide + Sulphuric acid.
5. Copper carbonate + Sulphuric acid.
6. Calcium carbonate + Hydrochloric acid.
7. Sodium hydroxide + Nitric acid.

1. Explain how you make copper sulphate
2. Explain the difference between an acid, alkali and a base
3. How do you measure pH?
4. Explain what happens when and acid and alkali react
5. Explain the difference between a strong and a weak acid
6. Explain the difference between a concentrated and dilute acid
7. Explain how weak acids react in water
8. Explain how strong acids react in water
9. Explain the relationship between concentration and pH

Complete the table

		Non-metals			
		fluorine	chlorine	oxygen	sulphur
Metals	lithium				
	sodium	NaF			
	potassium				
	magnesium				
	calcium				
	aluminium			Al_2O_3	