Metals

Metals make up almost 75% of the periodic table and are the familiar substances to man. Man-kind had used metals from thousands of years to make weapons, tools and from the last century, to conduct electricity.

They are the group I, II, the transition metals and some elements of group 3 and 4 and 5.

Properties of Metals:

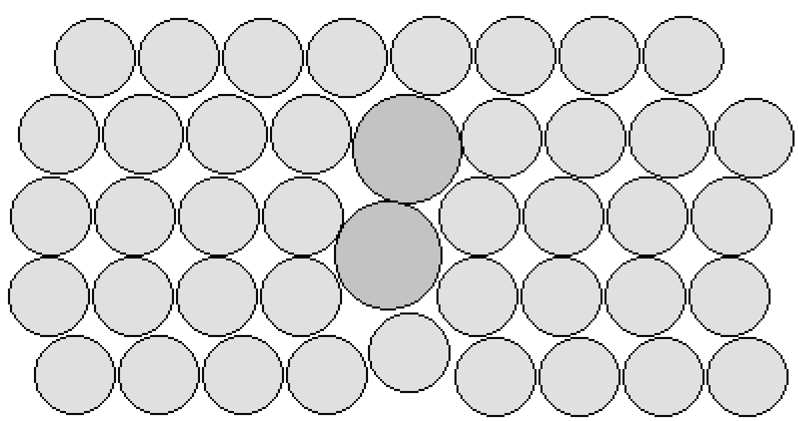
Physical Properties:

- Most metals are shiny and silvery in colour
- They are very good conductors of electricity due to free electrons in their structures that can carry the negative charge
- They are good conductors of heat due to the free electrons that can transmit heat throughout the structure
- They have high densities because of the ions closely packed in the lattice
- They have high melting and boiling points because of strong metallic bonds
- They are ductile (can be drawn into wires) and are malleable (can be beaten into sheets) because the layers of ions can move past each other without breaking the metallic bonds

Chemical Properties:

- They react with oxygen to form basic oxides
- Most react with dilute acid to make metal salt and hydrogen gas
- The more the reactive the metal, the higher its tendency to form compounds and vice-versa
- The more reactive a compound is the more stable its compounds are and vice-versa

Alloys:

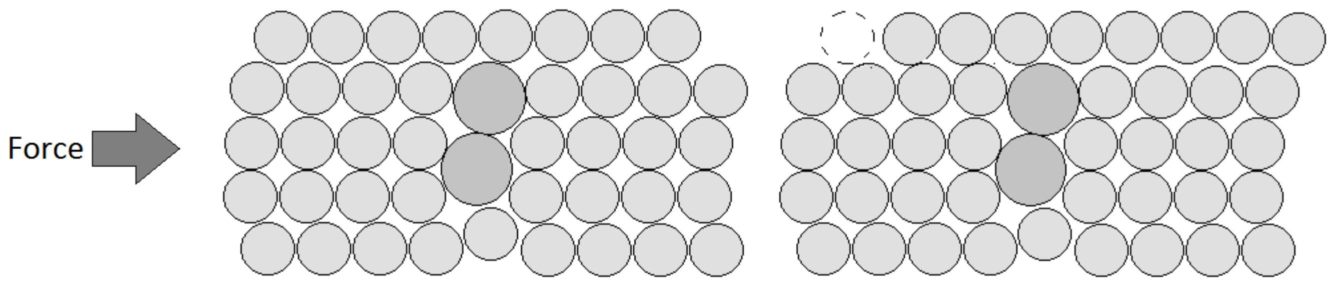


A pure metal doesn't have the exact properties that are needed for a particular job. Mixture of one or more metals called alloys can be created with properties that make them more suitable for specific purposes than pure metals.

An alloy is a mixture of metallic elements, and some non-metallic elements.

An alloy is made by weighing out correctly the different elements in different proportions and then mixing them together when they are molten.

An alloy is usually less ductile and malleable than pure metals. When force is applied to an alloy, the impurity atom stops the layers from slipping very far.



Some examples are:

Alloy	Typical composition		Particular properties
Brass	copper	70%	Harder than pure copper; gold coloured
	zinc	30%	
Bronze	copper	90%	Harder than pure copper
	tin	10%	
Mild Steel	iron	99.7%	Stronger and harder than pure iron
	carbon	00.3%	
Stainless Steel	iron	70%	Harder than pure iron; does not rust
	chromium	20%	
	nickel	10%	
Solder	tin	50%	lower melting point than either tin or lead
	lead	50%	

The reactivity series:

Reactivity series is a list of metals which are arranged in order of decreasing reactivity.

The more reactive the metal, the higher its tendency to form positive ions.

The more reactive the metal, the more stable are its compounds.

Potassium	K		Most Reactive
Sodium	Na		
Calcium	Ca		
Magnesium	Mg		
Aluminium	Al		
Carbon	C		
Zinc	Zn	Reactivity INCREASES upwards	
Iron	Fe		
Lead	Pb		
Hydrogen	H		
Copper	Cu		
Silver	Ag		
Gold	Au		Least Reactive

Metal reactions:

<table>
<tr><th>Metal</th><th>Symbol</th><th>Reaction when heated in air</th><th>Reaction with cold water</th><th>Reaction with dilute acids</th></tr>
<tr><td>Potassium</td><td>K</td><td rowspan="7">Burn to form oxides</td><td rowspan="3">React violently to produce OH and H_2</td><td rowspan="8">React to form metal salt and hydrogen gas</td></tr>
<tr><td>Sodium</td><td>Na</td></tr>
<tr><td>Calcium</td><td>Ca</td></tr>
<tr><td>Magnesium</td><td>Mg</td><td rowspan="4">React with steam to form basic oxides and H_2</td></tr>
<tr><td>Aluminium</td><td>Al</td></tr>
<tr><td>Zinc</td><td>Zn</td></tr>
<tr><td>Iron</td><td>Fe</td></tr>
<tr><td>Lead</td><td>Pb</td><td rowspan="2">Form oxides slowly (no burning)</td><td rowspan="4">NO REACTION</td></tr>
<tr><td>Copper</td><td>Cu</td><td rowspan="3">NO REACTION</td></tr>
<tr><td>Silver</td><td>Ag</td><td rowspan="2">NO REACTION</td></tr>
<tr><td>Gold</td><td>Au</td></tr>
</table>

Reduction of metal oxides with carbon:

Metals ABOVE carbon in the reactivity series cannot reduce their oxides.

Metals below carbon can be prepared by reduction of their oxides with carbon.

e.g.

$$Iron\ (III)\ Oxide\ +\ Carbon \rightarrow Iron\ +\ Carbon\ Dioxide$$

$$2\ Fe_2O_3\ +\ 3\ C \rightarrow 4\ Fe\ +\ 3\ CO_2$$

Reduction of metal oxides by a more reactive metal:

A metal oxide can reduced to the metal by heating the oxide with a more reactive metal.

e.g.

$$Iron\ (III)\ Oxide\ +\ Aluminium \rightarrow Iron\ +\ Aluminium\ Oxide$$

$$Fe_2O_3\ +\ 2\ Al \rightarrow 2\ Fe\ +\ Al_2O_3$$

Displacement reactions:

Any metal can displace a metal below it from its compounds.

If a piece of iron is put in copper (II) sulphate solution, a reaction takes place. The blue solution turns colourless and brown copper is deposited. Iron has displaced copper from its compound.

$$Iron\ +\ copper\ (II)\ sulphate \rightarrow Iron\ sulphate\ +\ Copper$$

$$Fe\ +\ CuSO_4 \rightarrow FeSO_4\ +\ Cu$$

Exception to the reactivity series:

Aluminium, although quite high in the reactivity series, takes time to react with an acid.

This is due to the layer of its oxide on it which acts like a protective layer, not allowing the acid to attack the underneath aluminium.

This layer acts in the same way as painting the metal, but is made from the metal itself.

Action of heat on Hydroxides of metals:

Metal	Hydroxide	Formula	Action of Heat
Potassium	Potassium Hydroxide	KOH	DON'T DECOMPOSE
Sodium	Sodium Hydroxide	$NaOH$	
Calcium	Calcium Hydroxide	$Ca(OH)_2$	Decompose to make metal oxide and oxygen
Magnesium	Magnesium Hydroxide	$Mg(OH)_2$	
Aluminium	Aluminium Hydroxide	$Al(OH)_3$	
Zinc	Zinc Hydroxide	$Zn(OH)_2$	
Iron	Iron Hydroxide	$Fe(OH)_3$	
Lead	Lead Hydroxide	$Pb(OH)_2$	
Copper	Copper Hydroxide	$Cu(OH)_2$	
Silver	Silver Hydroxide	$AgOH$	Decompose to the metal, water and oxygen

Action of heat on Metal nitrates:

Metal	Nitrate	Formula	Action of Heat
Potassium	Potassium Nitrate	KNO_3	Decompose to form metal NITRITE and oxygen
Sodium	Sodium Nitrate	$NaNO_3$	
Calcium	Calcium Nitrate	$Ca(NO_3)_2$	Decompose to make metal oxide, nitrogen dioxide and oxygen
Magnesium	Magnesium Nitrate	$Mg(NO_3)_2$	
Aluminium	Aluminium Nitrate	$Al(NO_3)_3$	
Zinc	Zinc Nitrate	$Zn(NO_3)_2$	
Iron	Iron Nitrate	$Fe(NO_3)_3$	
Lead	Lead Nitrate	$Pb(NO_3)_2$	
Copper	Copper Nitrate	$Cu(NO_3)_2$	
Silver	Silver Nitrate	$AgNO_3$	Decompose to the metal, nitrogen dioxide and oxygen

The extraction of metals:

Metals are widely used in everyday life. Very rarely are pure metals found free in the ground. Usually they are found in compounds with other materials in the form of ores.

Some very unreactive metals can be found alone in the earth's crust, like for example, gold. Other metals are found combined with other elements in the form of an ore.

Some ores include:

Metal	Ore	Chief chemical constituent
Aluminium	Bauxite	Aluminium Oxide Al_2O_3
Zinc	Zinc Blende	Zinc Sulphide ZnS
Iron	Haematite	Iron (III) Oxide Fe_2O_3

Before the metal is extracted from the ore, the ore is purified or concentrated.

The method used to extract the metal depends on its position in the reactivity series.

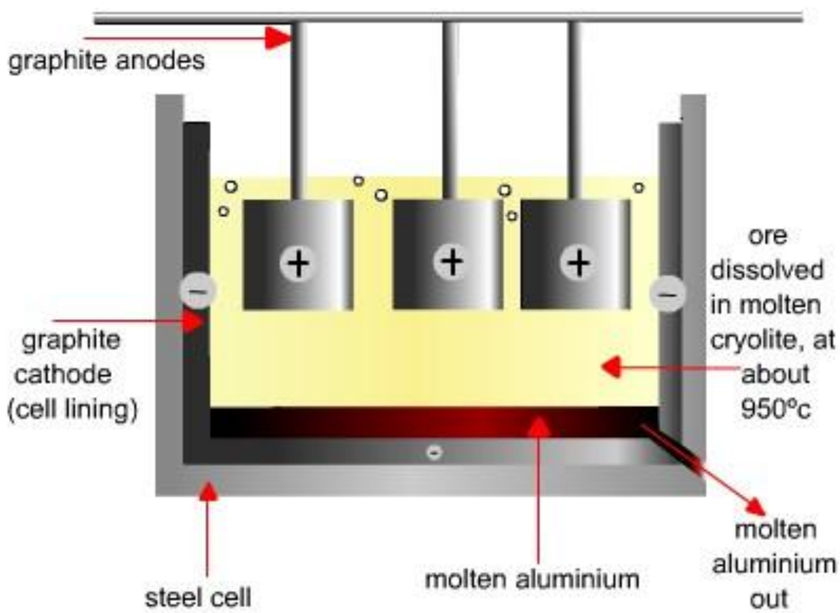
- Metals high in the reactivity series form stable ores, so they can only be got by electrolysis
- Metals in the middle of the reaction series form less stable ores, so carbon can be used to reduce the ore to the metal
- Metals low in the reactivity series are found uncombined in the earth's crust. If they are found in compounds, the compound should simply be heated, as it would decompose because the ore is unstable.

The extraction of Aluminium:

Aluminium is extracted from the ore BAUXITE or aluminium oxide Al_2O_3.

The aluminium oxide is INSOLUBLE so it is MELTED to allow the ions to move when the electric current is passed through it. Remember that electrolysis can only take place when the ions are MOLTEN or IN SOLUTION.

The anodes are made from CARBON and the cathode is a carbon-lined STEEL CASE.



At the cathode, aluminium metal forms:

$Al^{3+}{}_{(l)} + 3e^- = Al_{(l)}$

At the anode, oxygen is formed:

$2O^{2-}{}_{(l)} = O_{2(g)} + 4e^-$

The oxygen reacts with the carbon anodes and makes CO_2 gas. Anodes have to be replaced periodically.

Aluminium is extracted by electrolysis because it is high in the reactivity series and forms stable ores. It is also purified, before electrolysis, from other impurities like iron (III) oxide that can be found in the ore.

Aluminium oxide has a very high melting point, about 2000 degrees C, so it is dissolved in molten cryolite (Na_3AlF_6 Sodium Aluminium Fluoride) so its melting point is decreased from 2000 degrees C to about 900 degrees C.

This makes the aluminium ore able to be electrolysed at a lower cost (less heat has to be supplied to melt the ore).

The uses of aluminium:

Aluminium has low density and unlike iron, it resists corrosion due to a layer of its oxide on it, which prevents it from further oxidation.

It is used in the manufacture of:

- Air crafts; low density and strength
- Food containers; resistance to corrosion
- Overhead cables; good conductor of electricity and light
- Saucepans; good thermal conductivity
- Window frames; resistance to corrosion and low density

The extraction of iron:

Iron is extracted from haematite (Iron (III) Oxide) in a hot furnace. It is reduced by coke (carbon) because its ore is less stable (low in the reactivity series) and can be reduced using carbon.

Step 1:

Iron (III) oxide, coke and limestone are added into the hot furnace.

Step 2:

Hot air is pumped into the furnace. The coke burns in the hot air producing carbon dioxide and more heat, raising the temperature to about 1500 degrees C.

$$Carbon + Oxygen \rightarrow Carbon\ Dioxide$$

$$C + O_2 \rightarrow CO_2$$

Step 3:

The carbon dioxide is reduced by more carbon to form carbon monoxide.

$$Carbon\ Dioxide + Carbon \rightarrow Carbon\ Monoxide$$

$$CO_2 + C \rightarrow 2\ CO$$

Step 4:

The carbon monoxide reduces the iron (III) oxide into iron and carbon dioxide.

$$Iron\ (III) Oxide + Carbon\ Monoxide \rightarrow Iron + Carbon\ Dioxide$$

$$Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$$

Step 5:

The limestone (calcium carbonate) decomposes into calcium oxide and carbon dioxide because of heat.

$$Calcium\ Carbonate \rightarrow Calcium\ Oxide + Carbon\ Dioxide$$

$$CaCO_3 \rightarrow CaO + CO_2$$

Step 6:

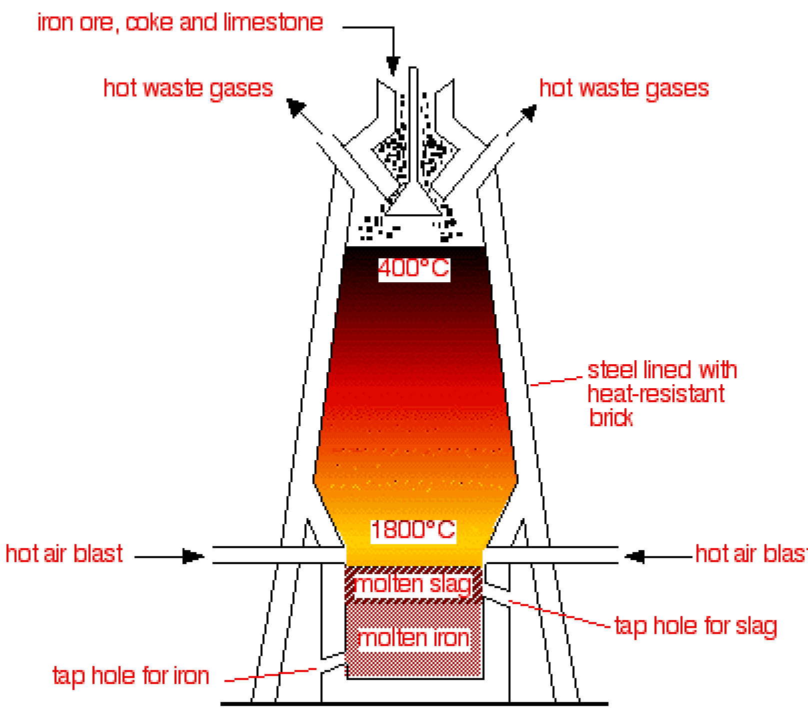
The impurities in the ore, like sand react with calcium oxide to make slag (Calcium Silicate).

$$Calcium\ Oxide + Silicon\ Dioxide \rightarrow Calcium\ Silicate$$

$$CaO + SiO_2 \rightarrow CaSiO_3$$

The molten iron sinks to the bottom of the furnace and the slag on the surface can taped off separately.

The slag is used in road making and in the manufacture of cement.



The iron the flows out of the furnace are called **cast iron**. It contains 4% carbon and other impurities like silicon, sulphur and phosphorous.

These lower the iron's melting point and make the iron expand slightly when cooled. The cast iron is also brittle.

To make the iron into steel, air or oxygen is bubbled through the molten cast iron. The impurities present are oxidised and leave as waste gases, but some like silica remain, which is removed by adding calcium oxide from decomposed limestone.

The slag formed is removed which floats over the molten iron.

Carbon can then be added by know amounts to make different types of steel, e.g. **mild steel** which is made of 0.25% carbon and the rest iron, and is used to make car bodies and machinery.

Other metal can be added to make different types of steel called alloys, e.g. **stainless steel**, which is made up of 18% chromium and 8% nickel, which is used to make cutlery and containers in chemical plants.

The extraction of zinc:

Zinc is found in the ore zinc Blende, which is zinc sulphide.

Step 1:

The ore is first concentrated. It is then roasted strongly in air to convert the zinc sulphide into zinc oxide. This called roasting.

$$Zinc\ Sulphide + Oxygen \rightarrow Zinc\ Oxide + Sulphur\ Dioxide$$

$$2\ ZnS + 3\ O_2 \rightarrow 2\ ZnO + 2\ SO_2$$

The sulphur dioxide is used in the manufacture of sulphuric acid.

Step 2:

The zinc oxide is then reduced by carbon in a blast furnace. The zinc formed is a gas.

$$Zinc\ Oxide + Carbon \rightarrow Zinc + Carbon\ Dioxide$$

$$2\ ZnO + C \rightarrow 2\ Zn + CO_2$$

Impurities are separated by the floatation method.

Depending on how much oxygen is present, carbon can form carbon monoxide. The zinc vapours pass out of the furnace and are cooled.

The uses of zinc:

- Galvanising steel; which is coating steel with a layer of zinc to protect it from rusting. Even if the layer is scratched the iron is still protected, because the zinc is more reactive than the iron.
- Making brass; an alloy of zinc 40% and copper 60%
- Making the outer coating of batteries

- **The uses of copper:**

Copper is used in the manufacture of:

- Brass; an alloy of copper 60% and zinc 40%
- Cooking utensils; good thermal conductivity
- Electrical wires; good conductor of electricity, ductile, unreactive and does not corrode.