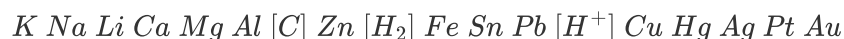


Chapter 1 - Metal Activity Series

Reactivity of Metals

- The **greater** a metal's **tendency** to form positive ions [*electropositivity*], the **more reactive** it is.

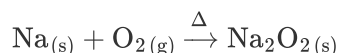
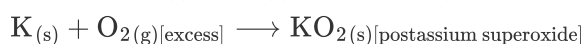
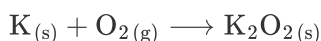
Reactivity Series



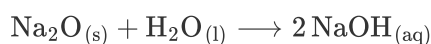
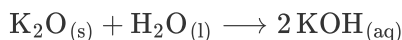
- Carbon is placed between aluminum and zinc because it can reduce zinc oxide to form zinc but cannot reduce aluminum oxide. Metals below carbon in the series can be extracted by heating their oxides with **carbon**, but those above carbon must be extracted by **electrolysis**.
- Same goes to hydrogen gas. It can reduce Iron(II/III) oxide to form iron but cannot reduce zinc oxide.
- The order of reactivity can be determined from experimental observations
 - reaction of metals with cold water or steam
 - reaction of metals with dilute acids

Reactions of Metals with Air (Oxygen)

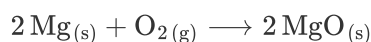
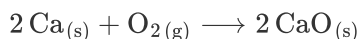
- Potassium and sodium** burn readily and **violently** with air to form **white peroxides**.



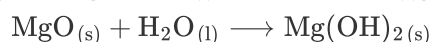
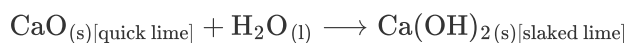
- Metals appear **dull** when exposed to air. Hence they have to be **stored under oil**.
- The oxides are **basic** in nature and **soluble** in water forming **alkalis** (soluble hydroxides).



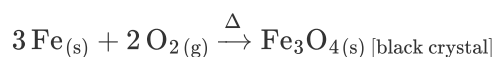
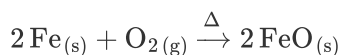
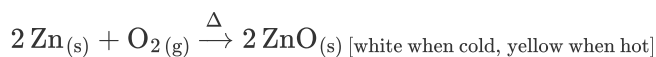
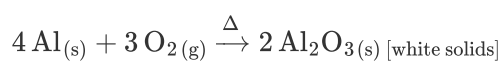
- Calcium and magnesium** burn **vigorously** in air to form **white oxides**.



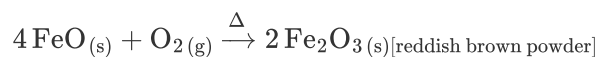
- Metals appear **dull** when exposed to air due to the layer of oxide on the surface.
- The oxides are **basic** in nature and **sparingly soluble** (low solubility) in water forming **hydroxides**.



- Aluminum, zinc and iron** also burn if they're **strongly heated** (particularly in oxygen)

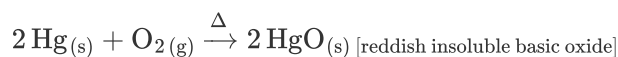
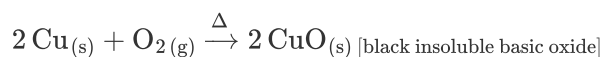
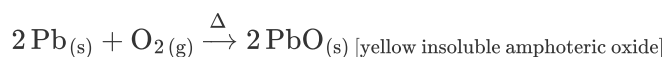
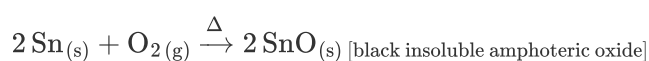


- The oxides of **aluminum and zinc** are amphoteric in nature and are insoluble in water.
 - **Amphoteric oxides** can react with both acids and alkalis to produce salts and water as the only products.
- The oxides layer of **aluminum and zinc** formed adheres to the surface and serves as a **protective layer** against further oxidation. [**Passivation**]
- Iron (II) oxide can be further oxidized to iron (III) oxide (under high temperature).



- The layer of oxide formed by **iron** is soft and falls off, exposing fresh metallic surface for **further oxidation**.

4. **Tin, lead, copper and mercury** have very little reaction with air. **Strong heating** is required and the **layer of oxide** normally forms on the **surface** only. (they **do not burn**)



5. **Silver and gold** have **no reaction** with air even on strong heating.

Reactions of Metals with Water

Reactions of Metals with Cold Water and Steam		
Metals	Reaction with cold water	Reactions with steam
K	$2\text{K}_{(s)} + 2\text{H}_2\text{O}_{(l)} \longrightarrow 2\text{KOH}_{(aq)} + \text{H}_{2(g)}$ <p>Reacts very vigorously and exothermically with cold water to form hydroxides and hydrogen gas. Enough heat is produced to cause hydrogen gas to catch fire and explode.</p>	Reacts explosively.
Na	$2\text{Na}_{(s)} + 2\text{H}_2\text{O}_{(l)} \longrightarrow 2\text{NaOH}_{(aq)} + \text{H}_{2(g)}$ <p>Reacts violently and exothermically with cold water to form hydroxides and hydrogen gas. Hydrogen gas may catch fire and explode.</p>	Reacts explosively.
Ca	$\text{Ca}_{(s)} + 2\text{H}_2\text{O}_{(l)} \longrightarrow \text{Ca}(\text{OH})_{2(aq)} + \text{H}_{2(g)}$ <p>Reacts with cold water but less vigorously than potassium and sodium.</p>	Reacts explosively.
Mg	$\text{Mg}_{(s)} + 2\text{H}_2\text{O}_{(l)} \longrightarrow \text{Mg}(\text{OH})_{2(s)} + \text{H}_{2(g)}$ <p>Reacts very slowly with cold water.</p>	$\text{Mg}_{(s)} + \text{H}_2\text{O}_{(g)} \xrightarrow{\Delta} \text{MgO}_{(s)} + \text{H}_{2(g)}$ <p>Hot magnesium reacts violently with steam to form manganese oxide (white solid) and hydrogen gas.</p>
Al	Aluminum doesn't react with cold water or steam as the layer of oxide on its surface serves as an inert layer of protection.	
Zn	No reactions occurs.	$\text{Zn}_{(s)} + \text{H}_2\text{O}_{(g)} \xrightarrow{\Delta} \text{ZnO}_{(s)} + \text{H}_{2(g)}$ <p>Hot zinc reacts readily with steam to produce zinc oxide and hydrogen gas. Zinc oxide is yellow when hot and white when cold.</p>
Fe	No reactions occurs.	$3\text{Fe}_{(s)} + 4\text{H}_2\text{O}_{(g)} \xrightarrow{\Delta} \text{Fe}_3\text{O}_{4(s)} + 4\text{H}_{2(g)}$ <p>Red hot iron react slowly with steam to form iron oxide and hydrogen gas.</p>
Sn & Pb	No reactions occurs.	Tin have lead have no reaction with steam at Bunsen burner temperature but show slight reaction at very high temperature.
Cu & Hg Ag & Au	No reactions occurs.	No reactions occurs.

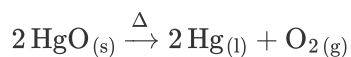
Reactions of Metals with Dilute Acids

Reactions of Metals with Dilute Hydrochloric Acid		
Metals	Observation	Equation
K & Na	React explosively with dilute acids and shouldn't be used in the laboratory.	$2\text{K}_{(s)} + 2\text{HCl}_{(aq)} \longrightarrow 2\text{KCl}_{(aq)} + \text{H}_{2(g)}$ $2\text{Na}_{(s)} + 2\text{HCl}_{(aq)} \longrightarrow 2\text{NaCl}_{(aq)} + \text{H}_{2(g)}$
Ca	Reacts vigorously, Effervescence of hydrogen gas. A colorless solution is formed.	$\text{Ca}_{(s)} + 2\text{HCl}_{(aq)} \longrightarrow \text{CaCl}_{2(aq)} + \text{H}_{2(g)}$
Mg	Reacts rapidly. Effervescence of hydrogen gas. A colorless solution is formed.	$\text{Mg}_{(s)} + 2\text{HCl}_{(aq)} \longrightarrow \text{MgCl}_{2(aq)} + \text{H}_{2(g)}$
Al	No reaction occurs because of the layer of aluminum oxide formed when the metal is exposed to moist air, Pure aluminum reacts and form hydrogen gas.	$2\text{Al}_{(s)} + 6\text{HCl}_{(aq)} \longrightarrow 2\text{AlCl}_{3(aq)} + 3\text{H}_{2(g)}$
Zn	Steady effervescence of hydrogen gas. A colorless solution is formed.	$\text{Zn}_{(s)} + 2\text{HCl}_{(aq)} \longrightarrow \text{ZnCl}_{2(aq)} + \text{H}_{2(g)}$
Fe	Slow effervescence of hydrogen gas. A pale green solution is formed.	$\text{Fe}_{(s)} + 2\text{HCl}_{(aq)} \longrightarrow \text{FeCl}_{2(aq)} + \text{H}_{2(g)}$
Sn	Very slow reaction with dilute HCl. Slightly soluble tin(II) chloride and hydrogen gas are formed.	$\text{Sn}_{(s)} + 2\text{HCl}_{(aq)} \longrightarrow \text{SnCl}_{2(aq)} + \text{H}_{2(g)}$
Pb	Reaction ceases almost immediately after the addition of diluted HCl because the insoluble lead(II) chloride formed prevents further reaction.	$\text{Pb}_{(s)} + 2\text{HCl}_{(aq)} \longrightarrow \text{PbCl}_{2(s)} + \text{H}_{2(g)}$
Cu & Hg Ag & Au	No reaction. Metals below hydrogen in the reactivity series cannot displace hydrogen gas from dilute acids.	

Thermal Decomposition of Metallic Compounds

Metal Oxides

- Most metal oxides are stable to heat except HgO and Ag_2O .



Metal Hydroxides

- Most metal hydroxides decompose into their **metallic oxides** and **water**.

Thermal Decomposition of Metal Hydroxides		
Metal hydroxide	Observation	Equation
Sodium hydroxide	White crystalline solid melts at 300°C without decomposing.	-
Calcium hydroxide	White powder decomposes. Liquid seen on the cooler side of the test tube. White residue left.	$\text{Ca(OH)}_2(\text{s}) \longrightarrow \text{CaO}(\text{s}) + \text{H}_2\text{O}(\text{l})$
Zinc hydroxide	White powder decomposes. Liquid seen on the cooler side of the test tube. Residue is yellow when hot and white when cold.	$\text{Zn(OH)}_2(\text{s}) \longrightarrow \text{ZnO}(\text{s}) + \text{H}_2\text{O}(\text{l})$
Iron(II) hydroxide	Dirty-green powder decomposes. Liquid seen on cooler side of the test tube. Residue is black when hot and turns reddish-brown when contact with air.	$\text{Fe(OH)}_2(\text{s}) \longrightarrow \text{FeO}(\text{s}) + \text{H}_2\text{O}(\text{l})$
Iron(III) hydroxide	Brown powder decomposes. Liquid seen on cooler side of the test tube. Residue is black when hot and brown when cold.	$2 \text{Fe(OH)}_3(\text{s}) \longrightarrow \text{Fe}_2\text{O}_3(\text{s}) + 3 \text{H}_2\text{O}(\text{l})$
Lead(II) hydroxide	White powder decomposes. Liquid seen on cooler side of the test tube. Residue is orange when hot and yellow when cold.	$\text{Pb(OH)}_2(\text{s}) \longrightarrow \text{PbO}(\text{s}) + \text{H}_2\text{O}(\text{l})$
Copper(II) hydroxide	Blue powder decomposes. Liquid seen on cooler side of the test tube. Black residue left.	$\text{Cu(OH)}_2(\text{s}) \longrightarrow \text{CuO}(\text{s}) + \text{H}_2\text{O}(\text{l})$

Metal Nitrates

- All metal nitrates decompose on heating. (They are best stored in an amber glass bottle or compatible opaque plastic container.)
- When metal nitrates(V) in solid form are strongly heated, they **decompose differently** according to their **thermal stability**.

Thermal Decomposition of Metal Nitrates		
Effect of heating	Nitrates	Observations
Decompose to give oxygen and their respective metal nitrites	Potassium nitrate Sodium nitrate	$2 \text{KNO}_3(\text{s}) \xrightarrow{\Delta} 2 \text{KNO}_2(\text{s}) + \text{O}_2(\text{g})$ $2 \text{NaNO}_3(\text{s}) \xrightarrow{\Delta} 2 \text{NaNO}_2(\text{s}) + \text{O}_2(\text{g})$ <p>White crystal melts to a colorless liquid, pale yellow solid left behind. Oxygen is evolved.</p>
	Calcium nitrate	$2 \text{Ca}(\text{NO}_3)_2(\text{s}) \xrightarrow{\Delta} 2 \text{CaO}(\text{s}) + 4 \text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ <p>White crystal melts, then decomposes to form a white residue. Brown fumes of nitrogen dioxide and oxygen is evolved.</p>
Decompose to give oxygen and brown fumes of nitrogen dioxide, leaving behind the respective metal oxides	Zinc nitrate	$2 \text{Zn}(\text{NO}_3)_2(\text{s}) \xrightarrow{\Delta} 2 \text{ZnO}(\text{s}) + 4 \text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ <p>White crystal melts, then decomposes to form a residue which is yellow when hot and white when cold, brown fumes of nitrogen dioxide and oxygen evolved.</p>
	Iron(II) nitrate	$2 \text{Fe}(\text{NO}_3)_2(\text{s}) \xrightarrow{\Delta} 2 \text{FeO}(\text{s}) + 4 \text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ <p>Green crystal melts, then decomposes to form a residue which is black when hot and turns reddish-brown on contact with air,</p>
	Lead(II) nitrate	$2 \text{Pb}(\text{NO}_3)_2(\text{s}) \xrightarrow{\Delta} 2 \text{PbO}(\text{s}) + 4 \text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ <p>White crystal decrepitates as it melts, then decomposes to form a residue which is orange when hot and yellow when cold. Brown fumes of nitrogen dioxide and oxygen is evolved.,</p>
	Copper(II) nitrate	$2 \text{Cu}(\text{NO}_3)_2(\text{s}) \xrightarrow{\Delta} 2 \text{CuO}(\text{s}) + 4 \text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ <p>Blue crystal melts, then decomposes to form a black residue. Brown fumes of nitrogen dioxide and oxygen evolved.</p>
Decomposes to give oxygen and brown fumes of nitrogen dioxide, leaving behind their respective metals	Mercury(II) nitrate	$2 \text{HgNO}_3(\text{s}) \xrightarrow{\Delta} 2 \text{Hg}(\text{l}) + 2 \text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ <p>White crystal melts, then decomposes to leave behind silvery mercury metal. Brown fumes of nitrogen dioxide and oxygen evolved.</p>
	Silver nitrate	$2 \text{AgNO}_3(\text{s}) \xrightarrow{\Delta} 2 \text{Ag}(\text{s}) + 2 \text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ <p>White crystal melts, then decomposes to leave behind silver metal. Brown fumes of nitrogen dioxide and oxygen evolved.</p>

Metal Carbonates

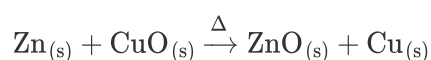
Thermal Decomposition of Metal Carbonates		
Effect of heating	Carbonate salt	Equation
Stable to heat, does not decompose	Potassium carbonate	-
	Sodium carbonate	-
Decomposes on heating to produce carbon dioxide and metallic oxides as residues.	Calcium carbonate	$\text{CaCO}_{3(s)} \longrightarrow \text{CaO}_{(s)} + \text{CO}_{2(g)}$
	Magnesium carbonate	$\text{MgCO}_{3(s)} \longrightarrow \text{MgO}_{(s)} + \text{CO}_{2(g)}$
	Aluminum carbonate	$\text{Al}_2(\text{CO}_3)_3(s) \longrightarrow \text{Al}_2\text{O}_{3(s)} + 3\text{CO}_{2(g)}$
	Zinc carbonate	$\text{ZnCO}_{3(s)} \longrightarrow \text{ZnO}_{(s)} + \text{CO}_{2(g)}$
	Iron(III) carbonate	$\text{Fe}_2(\text{CO}_3)_3(s) \longrightarrow \text{Fe}_2\text{O}_{3(s)} + 3\text{CO}_{2(g)}$
	Lead(II) carbonate	$\text{PbCO}_{3(s)} \longrightarrow \text{PbO}_{(s)} + \text{CO}_{2(g)}$
	Copper(II) carbonate	$\text{CuCO}_{3(s)} \longrightarrow \text{CuO}_{(s)} + \text{CO}_{2(g)}$
Decomposes on heating to produce carbon dioxide, oxygen and metal as residues.	Mercury(II) carbonate	$2\text{HgCO}_{3(s)} \longrightarrow 2\text{Hg}_{(l)} + 2\text{CO}_{2(g)} + \text{O}_{2(g)}$
	Silver carbonate	$2\text{Ag}_2\text{CO}_{3(s)} \longrightarrow 4\text{Ag}_{(s)} + 2\text{CO}_{2(g)} + \text{O}_{2(g)}$
	Gold carbonate	$2\text{AuCO}_{3(s)} \longrightarrow 2\text{Au}_{(s)} + 2\text{CO}_{2(g)} + \text{O}_{2(g)}$
Decomposes into ammonia, carbon dioxide and water without leaving any solid residue.	Ammonium carbonate	$(\text{NH}_4)_2\text{CO}_{3(s)} \longrightarrow 2\text{NH}_{3(g)} + \text{CO}_{2(g)} + \text{H}_2\text{O}_{(g)}$

Reduction of Metal Oxides

- A metal oxide may be reduced to the metal by:
 - reaction with other metals higher in the reactivity series.
 - reaction with non-metallic reducing metals e.g. carbon, carbon monoxide, hydrogen.
 - heating the oxide
- The method of reduction depends on the position of the metal in the reactivity series.

Reaction with other metals higher in reactivity series

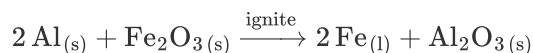
- A metal higher in the series can reduce the oxide of a metal that is lower in the series.
- e.g.



Thermit Reaction 铝热反应

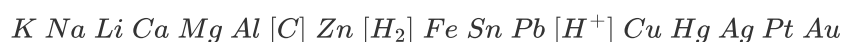
- When **aluminum powder** is used to **reduce the oxides of metals** lower in the reactivity series, the process is known as the thermit process. The reaction is **highly exothermic**.

e.g.



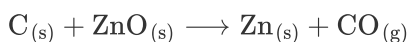
- The generated heat is so high that the **iron** obtained in the **molten state**. This process is often used in **welding**.
- In the laboratory, the process can be started by burning a piece of magnesium ribbon (as fuse 导火索) in the mixture of iron(III) oxide and aluminum. At the end of the reaction, a lump of solid iron will be obtained at the bottom of the crucible upon cooling. (<https://youtu.be/EDUwc953GOA?si=FtgYp70yrAUEswub>)

Reaction with non-metallic reducing agents



With Carbon

- Industrially, carbon in the form of coke (contains about 80%-85% carbon) is used to reduce zinc oxide to zinc.
- Zinc is the highest metal in reactivity series to be reduced by coke.



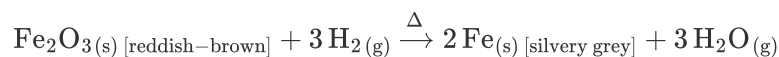
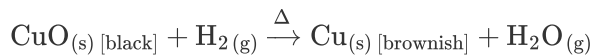
- Oxides of metals higher than carbon cannot be reduced by carbon.

The reactions of some metal oxides with carbon		
Metal oxide	Observations	Equation
Zinc oxide	White zinc oxide becomes yellow and glows when heated with carbon. Silvery zinc metal is formed.	$\text{ZnO}_{(s)} + \text{C}_{(s)} \xrightarrow{\Delta} \text{Zn}_{(s)} + \text{CO}_{(g)}$
Iron(III) oxide	Reddish-brown iron(III) oxide becomes black and glows when heated with carbon. Dull grey iron is formed.	$2 \text{Fe}_2\text{O}_{3(s)} + 3 \text{C}_{(s)} \xrightarrow{\Delta} 4 \text{Fe}_{(s)} + 3 \text{CO}_{2(g)}$
Tin(IV) oxide	White tin(IV) oxide glows when heated with carbon. Silvery tin metal is formed.	$\text{SnO}_{2(s)} + \text{C}_{(s)} \xrightarrow{\Delta} \text{Sn}_{(s)} + \text{CO}_{2(g)}$
Lead(II) oxide	Yellow lead(II) oxide glows when heated with carbon. Silvery beads of lead are formed.	$\text{PbO}_{2(s)} + \text{C}_{(s)} \xrightarrow{\Delta} \text{Pb}_{(s)} + \text{CO}_{2(g)}$
Copper(II) oxide	Black copper(II) oxide glows when heated with carbon. Reddish-brown copper metal is formed.	$2 \text{CuO}_{(s)} + \text{C}_{(s)} \xrightarrow{\Delta} 2 \text{Cu}_{(s)} + \text{CO}_{2(g)}$

With Hydrogen

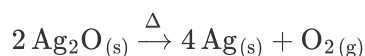
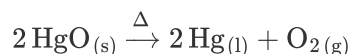
- At **Bunsen burner temperature**, hydrogen will reduce iron(III) oxide and oxides below it in the reactivity series

e.g.



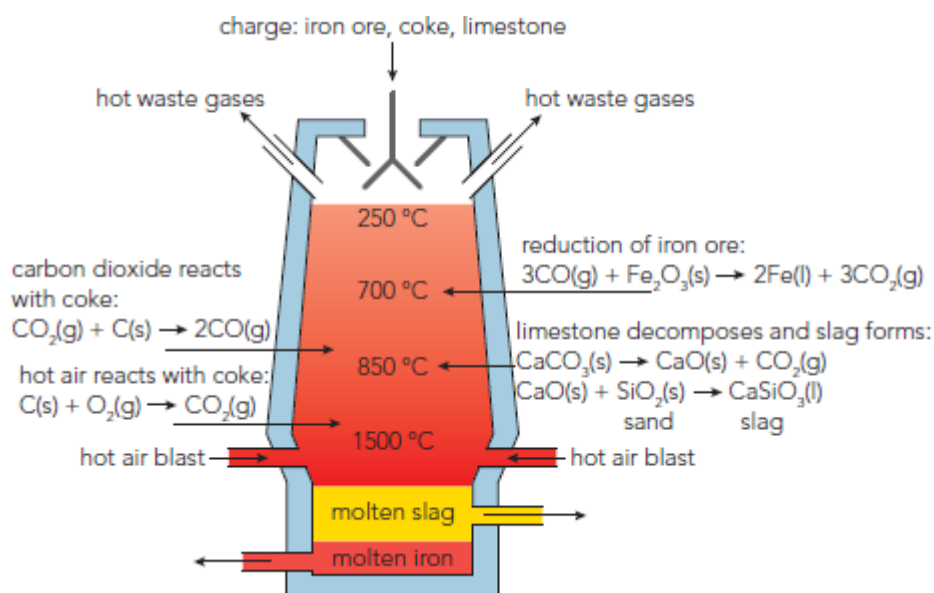
By heating the oxide

- Mercury(II) oxide and silver oxide are reduced by heating.



Extraction of Iron — The Blast Furnace

- Iron is extracted from its ore in the blast furnace.
- The main iron ore is called haematite. **Haematite** contains mainly **iron (III) oxide - Fe₂O₃**. The iron ore contains impurities, mainly **silica (silicon dioxide)**. **Limestone (calcium carbonate)** is added to the iron ore which reacts with the silica to form molten calcium silicate in the blast furnace. The **calcium silicate** (called **slag**) floats on the liquid iron. Since iron is below carbon in the reactivity series, iron in the ore is reduced to iron metal by heating with **carbon (coke)**. It is actually carbon monoxide which does the reducing in the blast furnace.

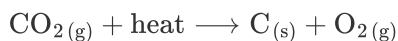
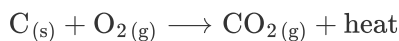


- The key features of the blast furnace extraction of iron from iron ore (hematite) are:
 - the burning of carbon (coke) to provide heat and produce carbon dioxide
 - the reduction of carbon dioxide to carbon monoxide
 - the reduction of iron(III) oxide by carbon monoxide
 - the thermal decomposition of limestone (calcium carbonate) to produce calcium oxide (lime)

- the use of limestone to remove the main impurity (sand) as slag (calcium silicate).

Chemistry of the Blast Furnace

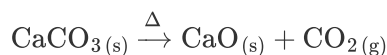
- **Hot air (oxygen)** is blasted into the furnace causing **coke (carbon)** to burn rapidly and raise the temperature to about 1800°C. The carbon dioxide then reacts with hot carbon to form **carbon monoxide**.



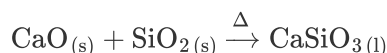
- Carbon monoxide then **reduces iron** in the ore (**haematite, magnetite**) to iron metal.



- The temperature where the reduction takes place is above 600°C, **Iron falls to the bottom** of the furnace where the temperature is 1800°C. Iron is **liquid** at this temperature and is tapped off periodically. Iron formed in blast furnace is **pig iron** which contains C (3.8-4.7%), Mn, Si, S and P.
- Melting point of pig iron is 1000-1300°C, much lower than pure iron, which is 1535°C.
- **Limestone** is calcium carbonate (CaCO₃) (as **flux**) and it is added to the blast furnace to **remove the impurities** in the iron ore. Calcium carbonate is decomposed by heat in the furnace to give **calcium oxide (quicklime)** and carbon dioxide.



- The **main impurity is silica (sand or rock)** which is **silicon dioxide**. Silicon dioxide is solid at the furnace temperature and the furnace would become blocked if it was not removed. Silicon dioxide reacts with calcium oxide to form **calcium silicate (slag)** which is **liquid** in the furnace. Slag flows to the bottom of the furnace where it **floats on the liquid iron** and is easily removed.



- **Slag layer protects the underlying iron from oxidation by the hot air blast.** Slag periodically tapped off from the iron and solidified, is used for road surfacing, cement manufacture, as insulating material.
- The **blast furnace gas** contains CO, CO₂, N₂ and etc.

*check out 讲义 page 11-13

anyway finally done yeahh. jiahuiiiiii @ 10th oct 2023 1745