Metals

Content

Reactivity of Metals

Metal	Extraction Method	Reaction with water (hydroxide produced)	Reaction with steam (oxide produced)	Reaction with HCl (metal chloride produced
Potassium K	Electrolysis	• Very violently. • Darts around surface of water with lilac flame. • Hydrogen gas explodes.	• Explosively • Darts around surface of water with lilac flame. • Hydrogen gas explodes.	• Explosively
Sodium Na		• Violently. • Darts around surface of water with yellow flame. • Hydrogen gas explodes.	• Explosively • Darts around surface of water with yellow flame. • Hydrogen gas explodes.	• Explosively
Calcium Ca		• Readily	• Explosively	• Violently
Magnesium Mg		• Slowly	• Violently • White glow	• Rapidly
Aluminium* Al		**NO REACTION**		
Carbon C				
Zinc* Zn	Reduced by carbon	**NO REACTION**	• Readily • ZnO formed is yellow	• Moderately fast
Iron Fe	Reduced by carbon or H_2		• Slowly	• Slowly
Lead Pb			**NO REACTION**	**NO REACTION**
Hydrogen H				
Copper Cu	Reduced by carbon or H_2	**NO REACTION**	**NO REACTION**	**NO REACTION**
Silver Ag				
Gold Au	Found uncombined			

Thermal stability of metal compounds

Metal	Metal Oxides	Metal Hydroxides	Metal Carbonates	Metal Nitrates
K	Stable to heating	Stable to heating	Stable to heating	Metal nitrite (NO_2) and oxygen
Na				
Ca		Metal oxide and steam	Metal oxide and carbon dioxide	Metal oxide, nitrogen dioxide and oxygen
Mg				
Al				
Zn				
Fe				
Sn				
Pb				
Cu				
Ag	Metal and oxygen	Do not form hydroxides	Metal, carbon dioxide and oxygen	Metal, nitrogen dioxide and oxygen
Au				

Recycling Metals

Benefits	Disadvantages
Conservation of metals	Processes during recycling create pollution (smelting)
Saves the need of extracting new metals, saving the environment as most mining and extraction processes damage the environment (pollution)	Costly: collecting and transporting scrap metals, sorting, separating and cleaning the scrap metals etc.
Prevent the landfills from filling up too quickly	Some alloys not recycled due to difficulty in separating metals in alloys

Extraction of Hematite

<u>Extraction of Iron (from Haematite in a Blast Furnace)</u>

1. Input: Haematite (which contains Iron(III) oxide Fe_2O_3), Coke ©, Limestone ($CaCO_3$) fed in from top.
2. Input: Blasts of hot air introduced at bottom.

3. Reactions:
   - Production of Carbon Dioxide: $C(s) + O_2(g) \rightarrow CO_2(g)$ + heat
   - Production of Carbon Monoxide: $CO_2(g) + C(s) \rightarrow 2CO(g)$
   - Production of Iron: $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$

   A small amount of Iron(III) Oxide reacts with Coke to give Iron and Carbon Monoxide.

4. Output: Waste gases (mainly Carbon Dioxide)
5. Reactions:
   - Thermal Decomposition of Limestone: $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
   - Neutralisation of Oxides: $SiO_2(s) + CaO(s) \rightarrow CaSiO_3(l)$

   Output: Molten slag $CaSiO_3(l)$
6. Output: Molten iron

<u>Extraction by Chemical Reduction</u>

When a metal is below Carbon in the reactivity series, the metal oxide can be reacted with coke. Carbon will displace the metal in the oxide, forming Carbon Dioxide.
eg. $2ZnO + C \rightarrow 2Zn + CO_2$

Hydrogen gas can reduce Lead(II) Oxide and Copper(II) Oxide to give the respective metal and Water. Under high temperature and pressure it can even reduce Iron(III) Oxide to give Iron and Water. However it is not usually used industrially as it is highly flammable.

Corrosion

Metal Corrosion

Corrosion is the gradual destruction of any metal due to reaction with air, water or other chemicals, forming cations by losing electrons.
In particular, **rusting** refers to the **corrosion of iron**, and needs both **water** and **air** to take place.

During rusting,
Iron atoms are oxidised to form Iron(II) ions: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
Oxygen and water accept electrons and are reduced: $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
Adding the half-equations gives us the overall reaction: $2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe^{2+}(aq) + 4OH^-(aq)$

In the presence of oxygen, Iron(II) ions can be oxidised to Iron(III) ions which react with hydroxide ions to form Iron(III) Hydroxide. $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
Iron(III) Hydroxide eventually changes into hydrated Iron(III) Oxide, which is rust.

Rusting is faster when :
1) **Water contains dissolved ionic substances**. This is because when the ionic compounds dissolve in water, the water effectively becomes an electrolyte which can **conduct electrons well**, hence speeding up corrosion.
2) **Iron is in contact with a less reactive metal**. Compared to a less reactive metal, Iron would have a **greater tendency to lose electrons** and form cations, and would hence lose the electrons to the less reactive metal on top of already losing electrons to oxygen and water, thus rusting more quickly. This is the working principle of a **simple electric cell**.

Rust Prevention

Rust Prevention

1) Protective layer - Prevents iron from being exposed. However if the layer is scratched off, exposed iron will start to rust.
2) Sacrificial Protection – As mentioned before, a more reactive metal would corrode while preventing a less reactive metal from corroding if the two are in contact, because of the difference in tendencies to lose electrons. By connecting iron to a more reactive metal, iron would be protected while the more reactive metal corrodes.
An example would be galvanizing, where iron is covered with a thin layer of zinc, as well as when magnesium or zinc blocks are attached to the sides of ships.

However, not all metals will corrode when exposed to air, as this depends on the nature of the metal oxide formed on the metal's surface.
In particular, aluminium oxide from the reaction of aluminium and oxygen forms a thin non-porous/ impervious layer on the metal's surface, preventing contact between oxygen, water and aluminium.
Meanwhile, iron oxide forms a porous layer on the metal's surface, so there is still contact between the oxygen, water and iron and thus further corrosion.
Further protection can be done by anodizing, where dyes are introduced and absorbed into the aluminium oxide layer, further preventing contact between oxygen, water and aluminium.

Alloys

Alloys

Pure metals tend to be malleable as the atoms are arranged in neat regular rows that slide across each other easily.
Alloys are a **mixture** of two or more substances (usually another metal or carbon).
Alloying a metal with other substances makes it harder and stronger as the **regular rows of metal atoms are disrupted by the presence of another type of atom** of different size, so the atoms can no longer slide across each other easily.

In particular, Steel is an important alloy made of iron and carbon, which is strong, elastic and tough.
During iron extraction in the blast furnace, the molten iron output solidifies as "pig iron" which contains about 5% carbon. 90% of this pig iron is converted into steel.
Oxygen is first blown into the molten iron to remove carbon impurities (reacting to form carbon dioxide).
The appropriate amounts of carbon and other additives are then added.

The properties of steel depend on its content of carbon and other metals, as well as heat treatment.
As carbon content increases, Hardness and Brittleness increase.

Alloy	**Composition**	**Advantages**	**Uses**
Mild Steel	99-99.5% Fe 0.15-0.25% C	Harder and stronger than iron; Can withstand great stress and strain	Construction of car bodies, machinery and steel rods to reinforce concrete.
Stainless Steel	90-95% Fe 5-10% Cr and Ni Variable % of C	Harder and stronger than iron; Resistant to corrosion; Very attractive in appearance	For making cutlery and surgical instruments; Used in chemical plants.