Study of the First Element: Hydrogen

Occurrence and Preparation of Hydrogen

Hydrogen is the first element in the periodic table. In elemental form, it exists as diatomic molecule H_2, and this H_2 is called dihydrogen. Electronic configuration of hydrogen is 1.

It was first discovered by Robert Boyle in 1672, though he was unable to establish that it was an element.

Hydrogen is placed separately in the periodic table. Do you know why?

It is due to its unique behavior. It contains only one electron. As a result, H^+ ion does not exist freely. It exists in association with other atoms or molecules.

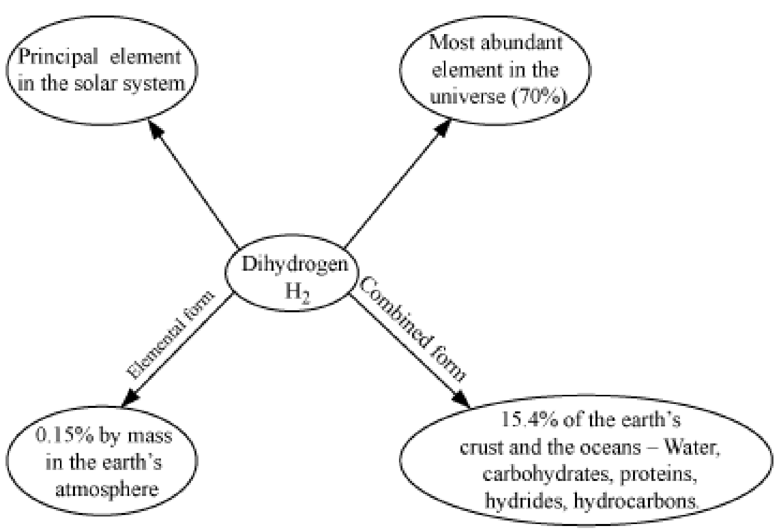
Curiosity Corner

Hydrogen is the lightest of all gases, and it burns with a pale-blue invisible flame.

Occurrence

As you probably know, hydrogen does not occur freely in nature; but in the sun and the other stars, free hydrogen is the chief constituent.

In combined state, you can find hydrogen occurring as water, in organic compounds, mineral products and acids.



Curiosity Corner

The large amount of energy in the form of heat and light is released by sun when two hydrogen atoms fuse to form a helium atom. This process of combination of two nuclei of hydrogen to form a bigger nucleus of helium is called nuclear fusion.

Similarities Between Hydrogen and Alkali metals

- One electron in the valence shell.
- Same valency (+1)
- Acts as reducing agents
- Burns in oxygen to form oxides

Similarities Between Hydrogen and Halogens

- One less electron than inert gas configuration
- Same valency
- Forms anions
- Electronegative in nature
- Same physical state
- Exists as diatomic molecule

Isotopes of hydrogen

There are three isotopes of hydrogen. They are

- Protium $({}_1^1H)$
- Deuterium $({}_1^2H \text{ or } D)$
- Tritium $({}_1^3H \text{ or } T)$

Preparation of hydrogen

Some methods for preparing hydrogen are given below.

- **By the action of dilute acids on metals:** Metals react with dilute sulphuric and hydrochloric acid to form their respective salts and hydrogen gas.

Metal + Acid (dilute) → Metallic salt + Hydrogen

$$Fe\ (s) + H_2SO_4\ (aq) \rightarrow FeSO_4\ (aq) + H_2\ (g)$$

$$Mg\ (s) + 2\,HCl\ (aq) \rightarrow MgCl_2\ (aq) + H_2\ (g)$$

- **By the action of water on metals**
- **Action of cold water:** Certain metals react with cold water to form their respective metallic hydroxides and hydrogen gas.

Metal + Water (Cold) → Metal Hydroxide + Hydrogen

$$2\,Na\ (s) + 2H_2O\ (l) \rightarrow 2\,NaOH\ (aq) + H_2\ (g)$$
$$2K\ (s) + 2H_2O\ (l) \rightarrow 2\,KOH\ (aq) + H_2\ (g)$$

$$Ca\ (s) + 2H_2O\ (l) \rightarrow Ca(OH)_2\ (aq) + H_2\ (g)$$

- **Action of hot water:** Certain metals do not react with cold metals but they react in hot water to form their respective metallic hydroxides and hydrogen gas.

Metal + Water (Hot) → Metal Hydroxide + Hydrogen

$$Mg\ (s) + 2H_2O\ (l) \rightarrow Mg(OH)_2\ (aq) + H_2\ (g)$$

- **Action of steam:** Certain metals react with steam to form their respective metallic oxides and hydrogen gas.

Metal + Steam → Metal Oxide + Hydrogen

$$2\,Al\ (s) + 3H_2O\ (g) \rightarrow Al_2O_3\ (s) + 3H_2\ (g)$$
$$Zn\ (s) + H_2O\ (g) \rightarrow ZnO\ (s) + H_2\ (g)$$
$$3\,Fe\ (s) + 4H_2O\ (g) \rightarrow Fe_3O_4\ (s) + 4H_2\ (g)$$

By the action of caustic alkalies on metals: Metals such as aluminium, zinc, and lead in powdered form, when boiled with concentrated solution of caustic soda (NaOH) or caustic potash (KOH), form their respective metallic salts and hydrogen gas.

Passing steam over red hot coke, produces water gas.

C (s) + H_2O (g) → CO (g) + H_2 (g)

$$\text{Metal} + \text{Caustic alkali solution} \xrightarrow{\text{boiling}} \text{Metal salt} + \text{Hydrogen}$$

$$Zn\,(s) + 2NaOH\,(aq) \xrightarrow{\text{boiling}} Na_2ZnO_2\ (aq) + 3H_2(g)$$

- Hydrogen is commercially prepared by the electrolysis of acidified water, using platinum electrodes.

$$2H_2O \xrightarrow[\text{Traces of acid/base}]{\text{Electrolysis}} 2H_2 + O_2$$

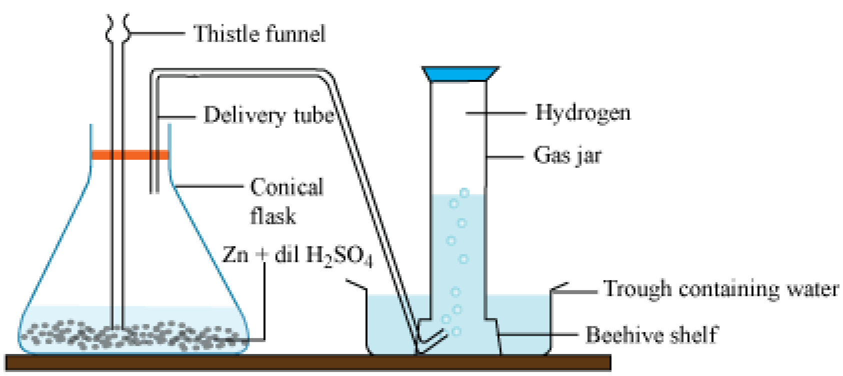
- **Laboratory preparation**

Hydrogen can be prepared in laboratories by the reaction of active metals like granulated zinc with dilute HCl or dilute sulphuric acid.

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$$

Let us see how.

Take 5 g zinc in a conical flask and fix a two-holed air tight stopper in the mouth of the flask. Through one hole, pass a thistle funnel, and through the other hole, pass a delivery tube whose other end is dipped in a water trough. Over the other end, place a beehive shelf.



Pour 2 cc of copper sulphate solution through the thistle funnel, and then dilute sulphuric acid, so that the lower end of the thistle funnel is completely immersed. Then, a reaction takes place between zinc and sulphuric acid, with the evolution of hydrogen gas. Here, the copper sulphate solution acts as a catalyst.

The evolved hydrogen gas bubbles through water. The first few bubbles contain air, so allow them to escape. Then, invert a gas jar completely filled with water over the beehive shelf. The hydrogen gas produced collects in the gas jar by displacing water.

The hydrogen gas prepared by this method may have the following impurities:

- Hydrogen sulphide (H_2S)
- Sulphur dioxide (SO_2)
- Oxides of nitrogen (NO_x)
- Phosphine (PH_3)
- Arsine (AsH_3)
- Carbon dioxide (CO_2)
- Water vapour (H_2O)

The following techniques are used to remove these impurities:

Type of Impurity	Technique	Reaction
Arsine and Phosphine	Using silver nitrate solution ($AgNO_3$)	$AsH_3 + 6\ AgNO_3 \rightarrow Ag_3As + 3\ AgNO_3 + 3\ HNO_3$ $PH_3 + 6\ AgNO_3 \rightarrow$ $Ag_3P + 3\ AgNO_3 + 3\ HNO3$
Hydrogen sulphide	Using lead nitrate solution ($Pb(NO_3)_2$)	$Pb(NO_3)_2 + H2S \rightarrow PbS + 2\ HNO_3$
Sulphur dioxide, Carbon dioxide, Oxides of nitrogen	Using caustic potash solution (KOH)	$SO_2 + 2\ KOH \rightarrow K_2SO_3 + H_2O$ $CO_2 + 2\ KOH \rightarrow K_2CO_3 + H_2$ $2\ NO_2 + 2\ KOH \rightarrow KNO_2 + KNO_3 + H_2O$
Water vapour	Using a suitable drying agent	For examples, fused calcium chloride ($CaCl_2$), caustic potash (KOH) stick, phosphorous pentoxide (P_2O_5)

- **Industrial method:**

  Bosch process

  Stage − I: Water gas is prepared by reacting coke with super-heated steam. The reaction is endothermic.

  $C + H_2O \rightarrow CO + H_2$

  Stage − II: Water gas is reacted with more of steam within a temperature range of 450 to 500^oC to form a mixture of carbon dioxide and hydrogen gas.

  $CO + H_2O + H_2 \rightarrow CO_2 + 2H_2$

  The catalyst used is Iron(III) oxide mixed with Chromium(III) oxide.

Stage −III: The primary impurities such as carbon monoxide are removed by absorbing them in ammoniacal copper (I) chloride.

$CO + CuCl + 2H_2O \rightarrow CO.CuCl.2H_2O$

Carbon dioxide is absorbed in caustic potash solution or dissolved in water at high pressures.

$CO_2 + 2KOH \rightarrow K_2CO_3 + H_2O$

$CO_2 + H_2O \rightarrow H_2CO_3$

Test for Hydrogen

1. Non-supporter of combustion
2. Burns in the air with pale blue flame forming water:

$2 H_2 + O2 \rightarrow 2 H_2O$

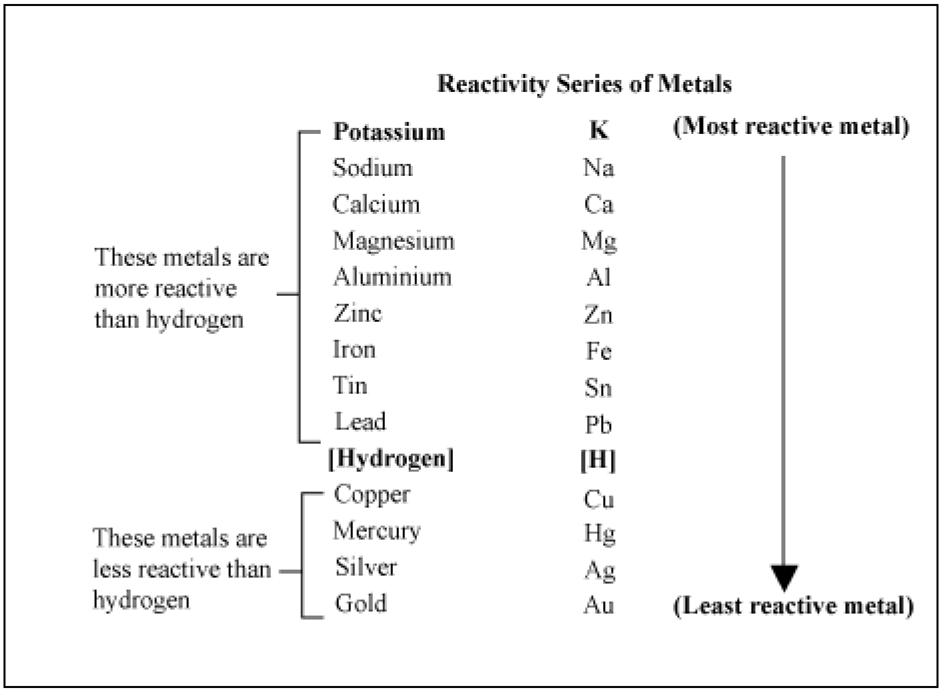
However, this test works only when hydrogen is not mixed with air (oxygen). If it is mixed with air(oxygen), then it burns with a popping sound.

Reactivity of Metals and Preparation of Hydrogen

The reactivity of metals can be determined by observing their reactions with salt solutions of other metals. When the reactivity of a metal is determined, it can be arranged in an increasing or decreasing order of their reactivities.

The series in which various metals are arranged in the order of their decreasing reactivity is called a Reactivity series.

This series is prepared by performing displacement reactions between various metals and their salt solutions. The reactivity series is given as follows:



In the reactivity series, metals present above a particular metal are more reactive than that metal, while the metals present below the particular metal are lesser reactive than it.

We know that all metals lose electrons to form positive ions. The tendency to lose electrons can be related to the reactivity of metals. If a metal can lose electrons easily, then it is very reactive. On the other hand, if its difficult for the metal to lose electrons, then it is less reactive.

Generally, we say that the metals present above **hydrogen** are more reactive than it, and displaces it from acids to liberate hydrogen gas. However, the metals present below hydrogen are less reactive than it, and cannot displace it from acids to liberate hydrogen gas.

Metals such as sodium and potassium (that lie above hydrogen) readily react with dilute acids to evolve hydrogen gas, whereas metals such as copper, gold, and silver (that lie below hydrogen) do not react with dilute acids.

Salient Features of Reactivity Series

- Metals are arranged in the decreasing order of their electropositive character.
- Metals at the top have greater reducing power. This power decreases, moving down the series.
- Metals at the top show greater tendency to get oxidised.
- Metals above hydrogen in the reactivity series liberate hydrogen gas from mineral acids.

- Metals at the top displace metals lower in the series from the aqueous solution of their salts.
- Metal oxides above Al, cannot be reduced by common reducing agents, the reverse is true for metal oxides below Al.
- Metals at the top having most electropositive nature combines with the most electronegative elements.

Displacement of Hydrogen from Dilute Acids

Metals react with hydrochloric acid in the similar fashion as they do with sulphuric acid. Sodium reacts very vigorously with hydrochloric acid to form a salt, and hydrogen gas is evolved in the reaction.

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\text{Hydrochloric acid}}{2HCl(aq)} \rightarrow \underset{\text{Sodium chloride}}{2NaCl(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Magnesium reacts vigorously with hydrochloric acid, but not as vigorously as sodium and potassium.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{Hydrochloric acid}}{2HCl(aq)} \rightarrow \underset{\text{Magnesium chloride}}{MgCl_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Zinc and iron also react with dilute hydrochloric acid to give zinc chloride and iron (II) chloride respectively. These reactions are comparatively less vigorous than the reaction of hydrochloric acid with aluminium metal.

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\text{Hydrochloric acid}}{2HCl(aq)} \rightarrow \underset{\text{Zinc chloride}}{ZnCl_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

$$\underset{\text{Iron}}{Fe(s)} + \underset{\text{Hydrochloric acid}}{2\,HCl(aq)} \rightarrow \underset{\text{Iron (II) Chloride}}{FeCl_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Thus, it can be concluded that metals react with acids to give a salt and hydrogen gas. The general equation for the process can be represented as:

$$\text{Metal} + \text{acid} \rightarrow \text{Salt} + \text{Hydrogen}$$

However, all metals do not react with dilute hydrochloric and sulphuric acids. Also, hydrogen gas is not evolved when a metal reacts with nitric acid. This is because nitric acid acts as an oxidising agent and oxidises hydrogen gas produced in the reaction to form water.

At the same time, nitric acid itself gets reduced to form nitrogen oxides such as nitrous oxide (N_2O), nitric oxide (NO), and nitrogen dioxide (NO_2). However, there are some metals such as magnesium, which react with very dilute nitric acid to evolve hydrogen gas.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{Nitric acid}}{2HNO_3(aq)} \rightarrow \underset{\text{Magnesium nitrate}}{Mg(NO_3)_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Metals such as gold and silver, which are very less reactive, do not react with acids. The only acid that dissolves gold is *aqua regia*. *Aqua regia* is the Latin name for ‘holy water’ or ‘royal water’.

It is called so because it is the only liquid that dissolves gold. It is prepared by mixing three parts of concentrated hydrochloric acid and one part of concentrated nitric acid. It is a highly corrosive and fuming solution having yellow or red colour. It can also dissolve platinum metal.

Displacement of Hydrogen by Alkalis

When metals react with base they forms hydrogen.

$$\begin{aligned}
&\text{Metal} + \text{Alkali} \rightarrow \text{Salt} + \text{Hydrogen}\\
&Zn + 2\ NaOH \rightarrow Na_2ZnO_2 + H_2\\
&Pb + 2\ NaOH \rightarrow Na_2PbO_2 + H_2\\
&2\ Al + 6\ NaOH \rightarrow 2\ Na_3AlO_3 + 3\ H_2\\
&Zn + 2\ KOH \rightarrow K_2ZnO_2 + H_2\\
&Pb + 2\ KOH \rightarrow K_2PbO_2 + H_2\\
&2\ Al + 6\ KOH \rightarrow 2\ K_3AlO_3 + 3\ H_2
\end{aligned}$$

K	These elements displace hydrogen from dil. acids (H_2SO_4 or HCl) with explosive violence
Na	
Ca	
Mg	These elements displace hydrogen from dil. acids (H_2SO_4 or HCl) vigorously
Al	
Zn	
F	These elements displace hydrogen from dil. acids (H_2SO_4 or HCl) gently
Ni	
Sn	
Pb	

H	
Cu	These elements do not displace hydrogen from dil. acids (H_2SO_4 or HCl)
Hg	
Ag	
Au	

Properties and Uses of Hydrogen

Let us first discuss the physical properties of hydrogen.

Physical Properties

- Pure hydrogen is a colourless, odourless, tasteless and combustible gas.
- Hydrogen is the lightest element. It is even lighter than air.

Curiosity Corner

Hydrogen was used for filling gas balloons. However, since it is highly explosive, balloons are now filled with helium.

- It is almost insoluble in water.
- It readily undergoes adsorption or occlusion with metals like palladium, platinum or nickel.

Chemical Properties

- Hydrogen is relatively inert at room temperature due to high H−H bond enthalpy.
- It is a combustible gas, but it does not support combustion.
- Hydrogen gas is neutral towards litmus paper.

Chemical reactions

Hydrogen undergoes chemical reactions by either of the following ways:

(i) Losing an electron to form H^+ ion

(ii) Gaining an electron to form H^- ion

(iii) Sharing an electron to form a single covalent bond

Let us see some of the reactions involving hydrogen.

- **Action with chlorine gas**

When a mixture of hydrogen and chlorine is exposed to sunlight, the mixture of gases turns colourless. This is because of the formation of hydrogen chloride gas.

$$H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$$

- **Action with oxygen**

Hydrogen burns in oxygen to form steam with pale blue flame, which condenses to form water.

$$2H_{2(g)} + O_{2(g)} \xrightarrow{\Delta} 2H_2O_{(steam)} + \text{Heat}$$

- **Action with nitrogen**

Hydrogen reacts with dinitrogen to form ammonia (Haber's process).

$$3H_{2(g)} + N_{2(g)} \xrightarrow[\text{Fe}]{\text{673K, 200atm}} 2NH_{3(g)};\ \Delta H^{\ominus} = -92.6\,\text{kJ mol}^{-1}$$

- **Action with metallic oxides**

When hydrogen gas is passed over heated oxides of metals, hydrogen reacts with them to form the corresponding metals along with water.

This reaction in which a metallic oxide loses its oxygen to form the pure metal is called **reduction** reaction. This property of hydrogen is thus called its **reducing** property.

$$CuO + H_2 \rightarrow Cu + H_2O$$
$$PbO + H_2 \rightarrow Pb + H_2O$$

- **Action with sulphur**

  When hydrogen gas is passed over molten sulphur, it reacts with it to form hydrogen sulphide gas.

  $H_2\,(g) + S_{(l)} \rightarrow H_2S_{(g)}$

- **Action with metals**

  When hydrogen gas is passed over heated metals, it reacts with them to form corresponding hydrides. These hydrides react with water to form corresponding metal hydroxides and liberate hydrogen gas.

$$2\,M + H_2 \xrightarrow{\Delta} 2 \underset{M = K,\ Na}{MH}$$

$$Ca + H_2 \xrightarrow{\Delta} CaH_2$$

- **Action with organic compounds**

  Hydrogen gas readily combines with unsaturated organic compounds that contain double and/or triple bond and converts them into corresponding saturated compounds. This reaction occurs in the presence of a catalyst like nickel metal and is known as **hydrogenation reaction.**

$$H_2C = CH_2 + H_2 \xrightarrow{Ni} H_3C - CH_3$$

$$HC \equiv CH + 2\,H_2 \xrightarrow{Ni} H_3C - CH_3$$

Uses of Hydrogen

- In the synthesis of ammonia gas
- In the manufacture of *vanaspati* fat by hydrogenation of *vanaspati* oils
- In chemical industries, for manufacturing compounds like methanol and hydrogen chloride $CO_{(g)} + 2H_{2(g)} \xrightarrow[\text{Catalyst}]{Co} CH_3OH_{(l)}$
- To reduce heavy metal oxides to the corresponding metals in metallurgical processes
- In atomic hydrogen and oxy-hydrogen torches, which are used for cutting and welding purposes
- As a rocket fuel
- In fuel cells to generate electricity
- In automatic lighters and self lighting gas jets.
- To produce artificial petrol from coal
- In extraction of metals by reducing the heated metallic oxides(less active metals) to metal

- In meteorological balloons, for studying weather conditions

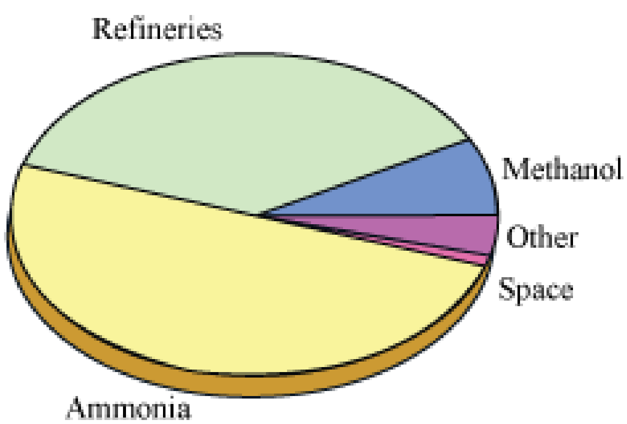
Curiosity Corner

Hydrogen is also used as a fuel in various cars like BMW, AUDI and CADILLAC.

It does not cause pollution and has more fuel efficiency than gasoline and other fuels.

Atomic hydrogen atoms recombine with the surface to be welded, raising the temperature to 4000 K

The following pi diagram shows the extent of use of hydrogen in different fields.



Redox Reactions: Oxidation and Reduction Reactions

Classical Idea of Redox Reactions

Oxidation

- Addition of oxygen or any other electronegative element or removal of hydrogen or any other electropositive element
- Examples:

Addition of oxygen

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(s)} + 2H_2O_{(l)}$$

Addition of electronegative element

$$Mg_{(s)} + Cl_{2(s)} \rightarrow MgCl_{2(s)}$$

$$Zn_{(s)} + S_{(s)} \rightarrow ZnS_{(s)}$$

Removal of hydrogen

$$2H_2S_{(g)} + O_{2(g)} \rightarrow 2S_{(s)} + 2H_2O_{(l)}$$

Removal of electropositive element

$$2K_4[Fe(CN)_6]_{(aq)} + H_2O_{2(aq)} \rightarrow 2K_3[Fe(CN)_6]_{(aq)} + 2KOH_{(aq)}$$

- Oxidising agents:
- Substances that oxidises other substance by accepting electrons or by providing oxygen or an electronegative ion, or by removing hydrogen or electropositive ion.
- In this process they reduce themselves.

State	**Oxidising agents**
Solid	Manganese dioxide (MnO_2), red lead, lead dioxide (PbO), potassium permanganate ($KMnO_4$), potassium dichromate ($K_2Cr_2O_7$), bleaching powder etc.
Liquid	Hydrogen peroxide (H_2O_2), concentrated nitric acid (HNO_3), concentrated sulphuric acid (H_2SO_4), bromine (Br_2) etc.
Gas	Oxygen, ozone (O_3), chlorine, etc.

- Tests for oxidising agents
- Strongly heating them releases oxygen which is good supporter of combustion and increases the flame of a burning splinter
- Bubbling hydrogen sulphide gas through the solution of oxidising agents leads to formation of a yellow precipitate of sulphur
- Warming them with concentrated hydrochloric acid releases chlorine gas which bleaches moist litmus paper
- Reacting them with acidified potassium iodide solution releases iodine gas; it turns freshly prepared starch solution blue

Reduction

- Addition of hydrogen or any other electropositive element or removal of oxygen or any other electronegative element
- Examples:

Addition of hydrogen

$CH_2{=}CH_{2(g)} + H_{2(g)} \rightarrow H_3C{-}CH_{3(g)}$

Addition of electropositive element

(Hg to $HgCl_2$)

$2HgCl_{2(aq)} + SnCl_{2(aq)} \rightarrow Hg_2Cl_{2(s)} + SnCl_{4(aq)}$

Removal of oxygen

$2HgO_{(s)} \xrightarrow{\Delta} 2Hg_{(l)} + O_{2(g)}$

Removal of electronegative element (Cl from $FeCl_3$) −−−

$2FeCl_{3(aq)} + H_{2(g)} \rightarrow 2FeCl_{2(aq)} + 2HCl_{(aq)}$

- Reducing agents:

- Substances that reduce other substances by providing electrons, or by providing hydrogen or an electropositive ion, or by removing oxygen or an electronegatiove ion.

- In this process they oxidise themselves.

State	Reducing agents
Solid	Carbon, metals such as zinc, aluminium, copper, sodium, stannous chloride ($SnCl_2$), glucose ($C_6H_{12}O_6$), etc.
Liquid	Hydrogen peroxide (H_2O_2), hydrogen iodide (HI), hydrogen bromide (HBr), etc.
Gas	Hydrogen sulphide (H_2S), carbon monoxide (CO), etc.

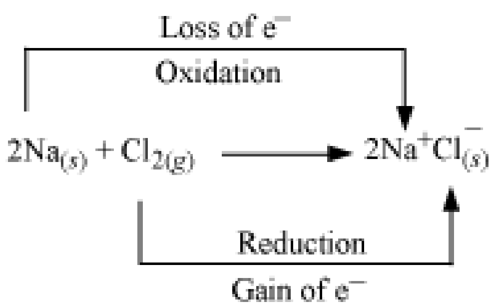
- Tests for reducing agents
- Heating them with black copper (II) oxide changes the oxide to red copper metal
- When warmed with nitric acid, they give out brown fumes of nitrogen dioxide.

- Adding dilute potassium permanganate solution to them decolourises the potassium permanganate solution
- Adding acidified potassium dichromate solution to them changes the colour of potassium dichromate solution from orange to green
- Adding iron (III) salts to them changes the colour of the salts from yellow to green (iron (II) salts are formed)

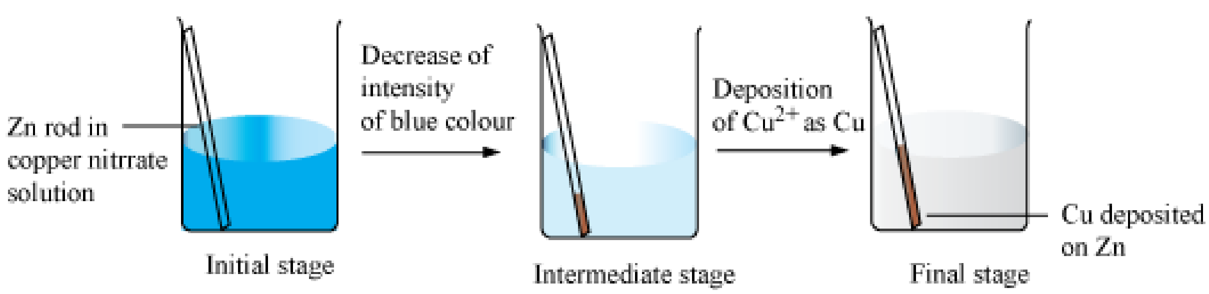
When oxidation and reduction occur simultaneously, such types of reactions are called **redox reactions**.

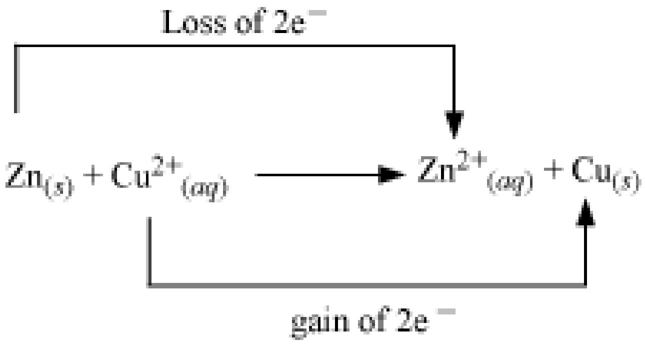
Redox Reactions in Terms of Electron Transfer Reactions

- Oxidation −−− Loss of electrons by any species
- Reduction −−− Gain of electrons by any species
- Oxidizing agent −Acceptor of electrons
- Reducing agent − Donor of Electrons
- Examples:



- Competitive electron transfer reactions
- Reaction between metallic zinc and the aqueous solution of copper nitrate:





- Reaction between metallic copper and the aqueous solution of zinc sulphate:

  $Cu_{(s)} + Zn^{2+}_{(aq)} \rightarrow$ No reaction

  That is, Zn has greater tendency to lose electrons than Cu.

- Metal activity series or electrochemical series:

$$\xrightarrow[\text{Decreasing order of tendency to lose electrons}]{K > Na > Li > Mg > Al > Zn > Fe > Co > Cu > Ag > Au}$$

Differences between oxidation and reduction

Oxidation	Reduction
It is addition of oxygen	It is removal of oxygen
It is removal of hydrogen	It is addition of hydrogen
It is addition of an electronegative atom/ion	It is removal of an electronegative atom/ion
It is removal of electropositive atom/ion	It is addition of electropositive atom/ion
There is an increase in positive valency	There is a decrease in positive valency
There is a decrease in negative valency	There is an increase in negative valency
Loss of electrons occur	Gain of electrons occur