UNIT-7

OXIDATION AND REDUCTION

Oxidation: oxygen is gained

Magnesium burns in air with a dazzling white flame.
A white ash is formed. The reaction is:

magnesium + oxygen → magnesium oxide

$$2Mg\ (s) + O_2\ (g) \rightarrow 2MgO\ (s)$$

The magnesium has gained oxygen.
We say it has been oxidised.

A gain of oxygen is called oxidation.
The substance has been oxidised.

Oxidation and Reduction

An **oxidation** reaction is one in which **oxygen is added** to a substance.

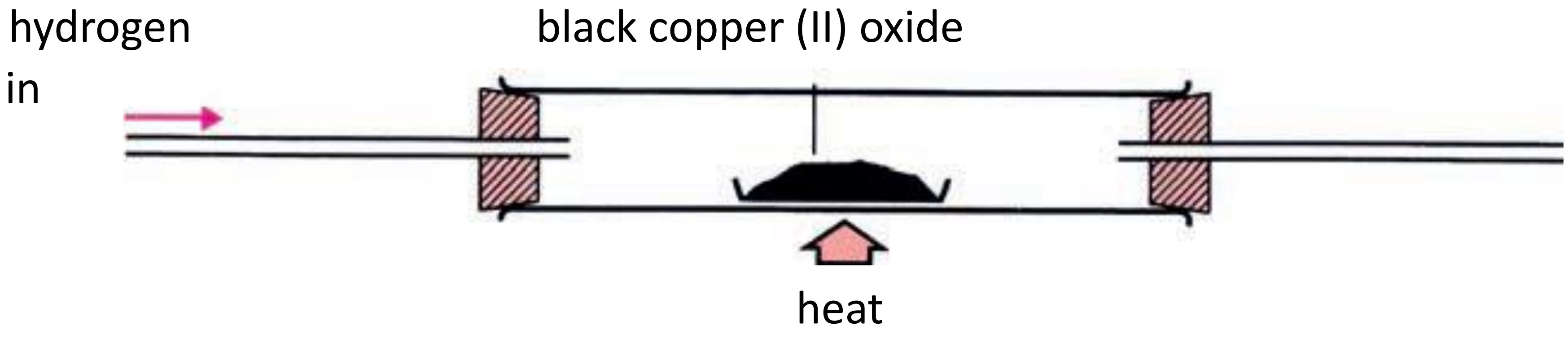
Example:

Methane is oxidised when it burns in air. Oxygen is added to the carbon in methane, forming carbon dioxide. Oxygen is also added to the hydrogen in methane, forming water.

$$\underset{\text{methane}}{CH_{4(g)}} + \underset{\text{oxygen}}{2\,O_{2(g)}} \rightarrow \underset{\text{carbon dioxide}}{CO_{2(g)}} + \underset{\text{water}}{2\,H_2O_{(l)}}$$

Reduction: oxygen is lost

Now look what happens when hydrogen is passed over heated copper(II) oxide. The black compound turns pink:



This reaction is taking place:

copper(II) oxide + hydrogen → copper + water

$CuO\ (s) + H_2\ (g) \rightarrow Cu\ (s) + H_2O\ (l)$

This time the heated substance is losing oxygen. It is being reduced.

A loss of oxygen is called reduction. The substance is reduced.

Oxidation and reduction take place together

Look again at the reaction between copper(II) oxide and hydrogen.

Copper(II) oxide loses oxygen, and hydrogen gains oxygen:

oxidation

$$CuO\ (s) + H_2\ (g) \rightarrow Cu\ (s) + H_2O\ (l)$$

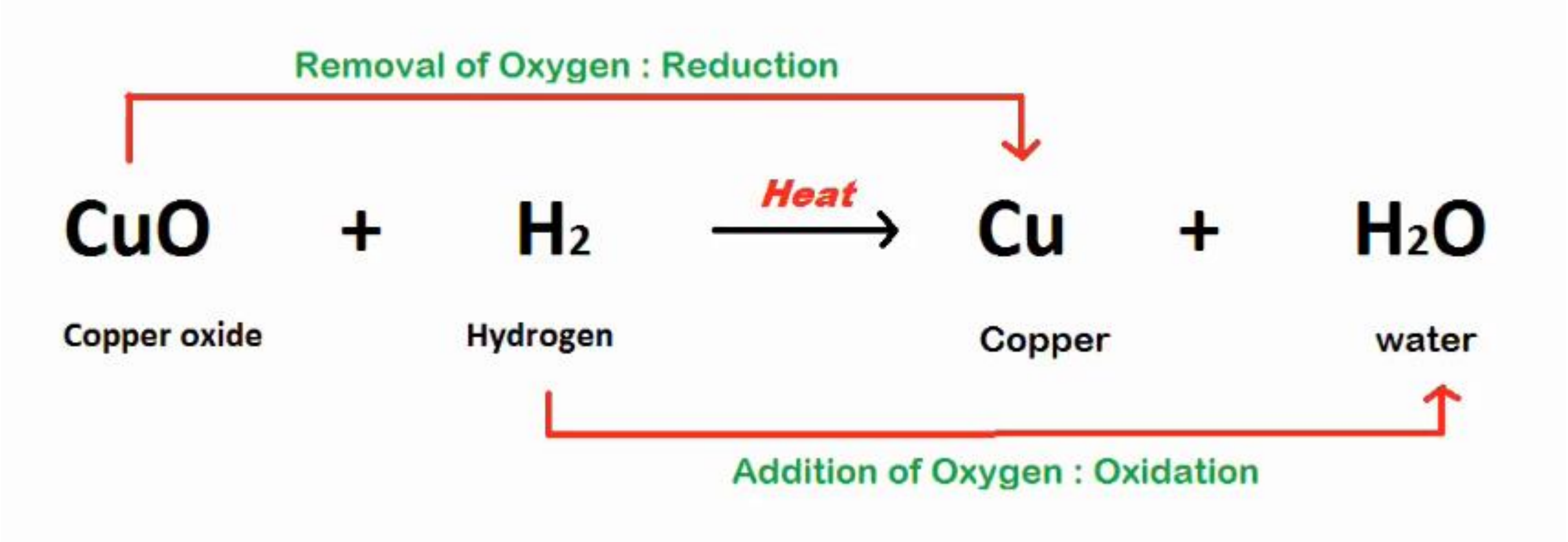
reduction

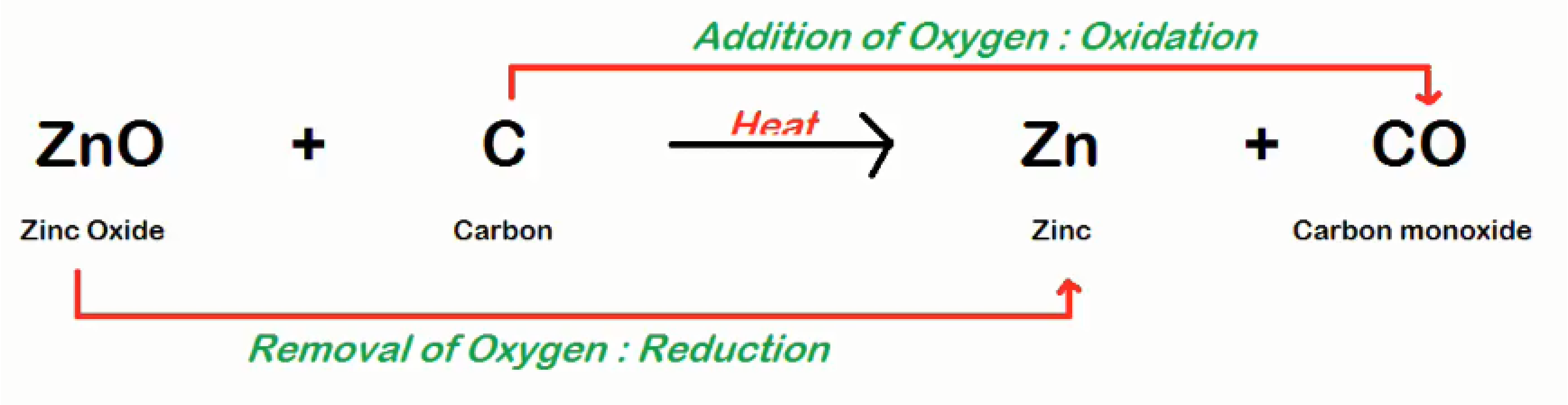
So the copper(II) oxide is reduced, and the hydrogen is oxidised.

Oxidation and reduction always take place together.

So the reaction is called a **Redox Reaction**.

Oxidation and Reduction Examples





Few more examples of redox reactions

The reaction between calcium and oxygen:

oxidation

$$2Ca\,(s) + O_2\,(g) \rightarrow 2CaO\,(s)$$

reduction

The reaction between hydrogen and oxygen:

oxidation

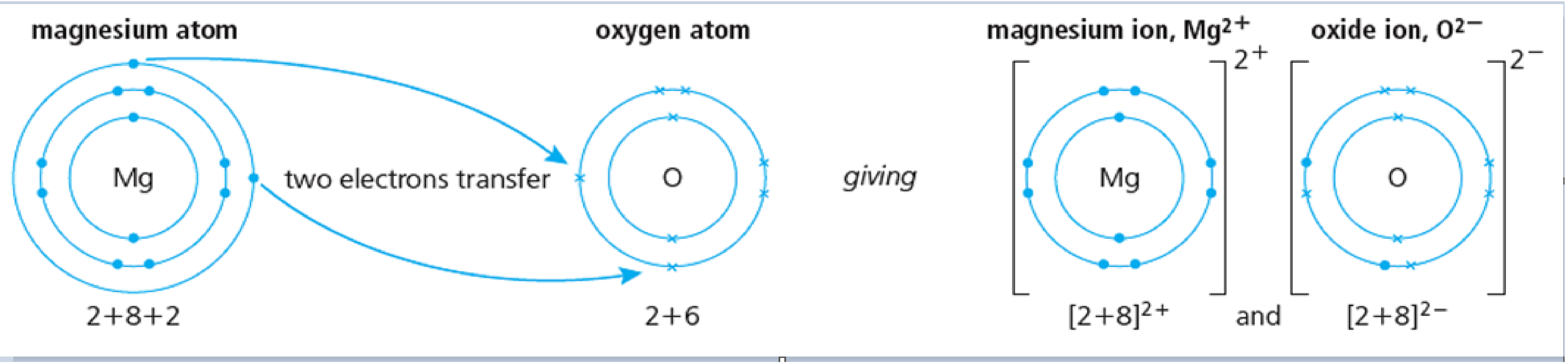
$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l)$$

reduction

Redox and electron transfer

When magnesium burns in oxygen, magnesium oxide is formed:

$$2Mg\ (s) + O_2\ (g) \rightarrow 2MgO\ (s)$$



Writing half-equations to show the electron transfer

magnesium: $2\ Mg \rightarrow 2\ Mg^{2+} + 4e^-$ $\rightarrow$ OXIDATION

oxygen: $O_2 + 4e^- \rightarrow 2O^{2-}$ $\rightarrow$ REDUCTION

Redox without oxygen

Example 1 : The reaction between sodium and chlorine

The equation is:

$$2Na\ (s) + Cl_2\ (g) \rightarrow 2NaCl\ (s)$$

Half equations:

sodium: $2Na \rightarrow 2Na^+ + 2e^-$ (oxidation)

chlorine: $Cl_2 + 2e^- \rightarrow 2Cl^-$ (reduction)

The reaction is redox reaction.

Example 2 : The reaction between chlorine and potassium bromide

$$Cl_2\ (g) + 2KBr\ (aq) \rightarrow 2KCl\ (aq) + Br_2\ (aq)$$

colourless orange

$[2K^+Br^-]$ $[2K^+Cl^-]$

Half equations:

chlorine: $Cl_2 + 2e^- \rightarrow 2Cl^-$ (reduction)

bromide ion: $2Br^- \rightarrow Br_2 + 2e^-$ (oxidation)

The reaction is redox reaction.

Oxidation states

What does oxidation state mean?

Oxidation state tells you how many electrons each atom of an element has gained, lost, or shared, in forming a compound.

An oxidation state is also called oxidation number.

An oxidation number that is assigned to an atom in a substance.
The oxidation number could be +ve, -ve or zero. It indicates whether electrons are lost or gained.

- ✓ If the oxidation number is **+ve**, this means that atom **looses electrons**.
- ✓ If the oxidation number is **–ve**, this means that atom **gains electrons**.
- ✓ If the oxidation number is **zero**, this means that atom **neither gains nor looses electrons.**

The rules for oxidation states

$$\underset{0}{2Na\,(s)} + \underset{0}{Cl_2\,(g)} \longrightarrow \underset{+\mathrm{I}\;-\mathrm{I}}{2NaCl\,(s)}$$

1 Each atom in a formula has an oxidation state.

2 The oxidation state is usually given as a Roman numeral.
Note these Roman numerals:

number	0	1	2	3	4	5	6	7
Roman numeral	0	I	II	III	IV	V	VI	VII

3 Where an element is not combined with other elements, its atoms are in oxidation state 0.

4 Many elements have the same oxidation state in most or all their compounds. Look at these:

Element	Usual oxidation state in compounds
hydrogen	+I
sodium and the other Group I metals	+I
calcium and the other Group II metals	+II
aluminium	+III
chlorine and the other Group VII non-metals, in compounds without oxygen	−I
oxygen (except in peroxides)	−II

5 But atoms of transition elements can have variable oxidation states in their compounds. Look at these:

Element	**Common oxidation states in compounds**
iron	+II and +III
copper	+I and +II
manganese	+II, +IV, and +VII
chromium	+III and +VI

So for these elements, the oxidation state is included in the compound's name. For example iron(III) chloride, copper(II) oxide.

6 Note that in any formula, the oxidation states must add up to zero. Look at the formula for magnesium chloride, for example:

$$MgCl_2$$

$$+\mathrm{II} \qquad 2 \times -\mathrm{I} = -\mathrm{II} \quad \text{Total} = \text{zero}$$

So you could use oxidation states to check that formulae are correct.

Common Oxidation numbers

All elements in uncombined state	0
All monoatomic ions	Same as their charge
Hydrogen in all compounds	+1
Hydrogen in hydrides(NaH,MgH_2)	-1
Oxygen in all compounds	-2
Oxygen in peroxides (H_2O_2)	-1
All group 1 elements in compounds	+1
All group 2 elements in compounds	+2
All halogens in group 1 and 2 compounds	-1
The sum of oxidation numbers in a compound	0
The sum of oxidation numbers in a polyatomic ion	Same as its charge

EXAMPLE -1

What is the oxidation number of sulfur in sodium sulfate Na_2SO_4?

Sodium oxidation number=+1

Sulfur oxidation umber=x

Oxygen oxidation number=-2

$2(+1)+x+4(-2)=0$

$2+x-8=0$

$X-6=0$

$x=+6$

EXAMPLE -2

What is the oxidation number of manganese in a permanganate ion, MnO_4^-?

Manganese oxidation number=x

Oxygen oxidation number=-2

$X+4(-2)=-1$ (Hint: total of oxidation number is equal to charge of the ion)

$X-8=-1$

$X=+7$

EXAMPLE -3

Determine the oxidation number of S in SO_2

$$X+2(-2)=0$$

$$X-4=0$$

$$X=+4$$

EXAMPLE -4

Determine the oxidation number of S in SO_4^{2-}

$$X+4(-2)=-2$$

$$X-8=-2$$

$$X=+6$$

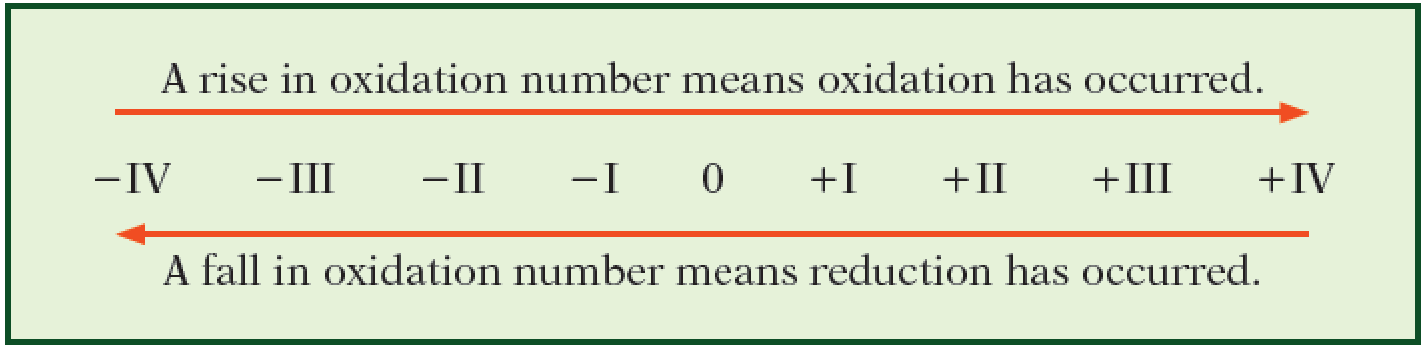
Oxidation states change during redox reactions

Look at the equation for the reaction between sodium and chlorine:

$$2Na\ (s) + Cl_2\ (g) \rightarrow 2NaCl\ (s)$$

$$0 \qquad\quad 0 \qquad\qquad\quad +I \quad -I$$

If oxidation states change during a reaction, it is a redox reaction.



Using oxidation states to identify redox reactions

Example 1 Iron reacts with sulfur to form iron(II) sulfide:

$$\underset{0}{Fe\,(s)} + \underset{0}{S\,(s)} \longrightarrow \underset{+II\ -II}{FeS\,(s)}$$

There has been a change in oxidation states. So this is a redox reaction.

Example 2 When chlorine is bubbled through a solution of iron(II) chloride, iron(III) choride is formed. The equation and oxidation states are:

$$\underset{+II\ -I}{2FeCl_2\,(aq)} + \underset{0}{Cl_2\,(aq)} \longrightarrow \underset{+III\ -I}{2FeCl_3\,(aq)}$$

There has been a change in oxidation states. So this is a redox reaction.

Example 3 When ammonia and hydrogen chloride gases mix, they react to form ammonium chloride. The equation and oxidation states are:

$$\underset{-III\ +I}{NH_3\,(g)} + \underset{+I\ -I}{HCl\,(g)} \longrightarrow \underset{-III\ +I\ -I}{NH_4Cl\,(s)}$$

There has been no change in oxidation states. So this is *not* a redox reaction.

Oxidising and Reducing Agents

When hydrogen reacts with heated copper(II) oxide, the reaction is:

$$\text{copper(II) oxide} + \text{hydrogen} \longrightarrow \text{copper} + \text{water}$$
$$CuO\ (s) + H_2\ (g) \longrightarrow Cu\ (s) + H_2O\ (l)$$

The copper(II) oxide is **reduced** to copper by reaction with hydrogen.
So hydrogen acts as a **reducing agent**.

The hydrogen is itself **oxidised** to water, in the reaction. So copper(II) oxide acts as an **oxidising agent**.

An oxidising agent oxidises another substance – and is itself reduced.
A reducing agent reduces another substance – and is itself oxidised.

Examples for Oxidising and Reducing Agents

Reduction: CuO is reduced;
CuO is the oxidizing agent.

$$CuO + H_2 \longrightarrow Cu + H_2O$$

Oxidation: H_2 is oxidized;
H_2 is the reducing agent.

Oxidising and reducing agents in the lab

Some substances have a strong drive to gain electrons. So they are strong oxidising agents. They readily oxidise other substances by taking electrons from them. Examples are oxygen and chlorine.

Some substances are strong reducing agents, readily giving up electrons to other substances. Examples are hydrogen, carbon monoxide, and reactive metals like sodium.

Some oxidising and reducing agents show a colour change when they react. This makes them useful in lab tests. Look at these three examples.

Example:

1 **Potassium manganate(VII): an oxidising agent**

2 **Potassium dichromate(VI): an oxidising agent**

3 **Potassium iodide: a reducing agent**