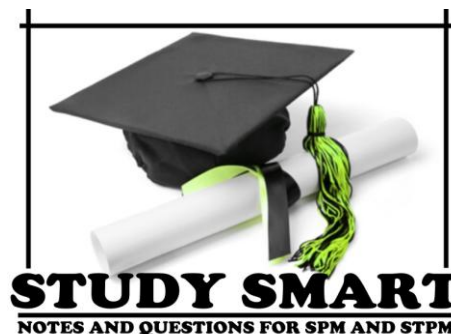


STUDYSMART  
CHEMISTRY FORM 5  
CHAPTER 3 : OXIDATION AND REDUCTION

- 3.1 Analysing Redox Reactions
- 3.2 Analysing rusting as a redox reaction
- 3.3 Understanding the reactivity series of metals and its application
- 3.4 Analysing redox reactions in electrolytic and chemical cells
- 3.5 Appreciating the ability of elements to change their oxidation numbers



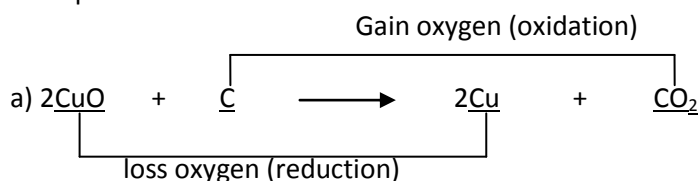
### 3.1 ANALYSING REDOX REACTIONS

- Redox reactions are chemical reaction involving oxidation and reduction occurring simultaneously

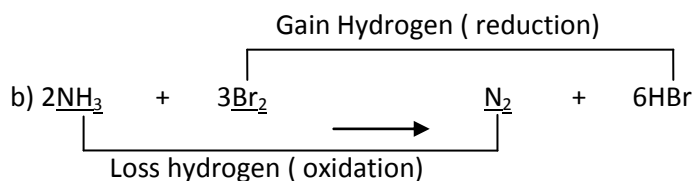
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Process	Oxidation	Reduction
In term of loss of or gain oxygen	Substance gain oxygen	Substance loss oxygen
Loss or gain hydrogen	Substance loss hydrogen	Substance gain hydrogen
Transfer of electron	Substance loss electron	Substance gain electron
Changes in oxidation number	Oxidation number increases	Oxidation number decreases

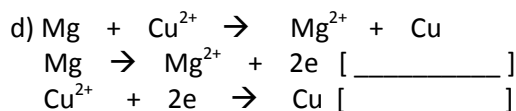
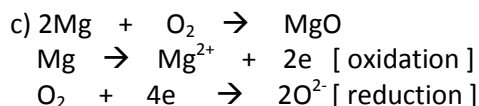
- Examples



- \* Oxidized substance : Carbon C
- \* Reduction substance : Copper(II)oxide, CuO
- \* Oxidizing agent : Copper(II)oxide, CuO
- \* Reducing agent : Carbon C



- \* Oxidized substance : Ammonia, NH<sub>3</sub>
- \* Reduction substance : Bromine, Br
- \* Oxidizing agent : Bromine, Br
- \* Reducing agent : Ammonia, NH<sub>3</sub>



### Oxidation Number

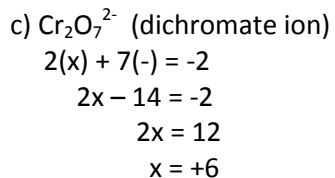
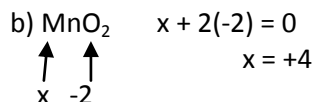
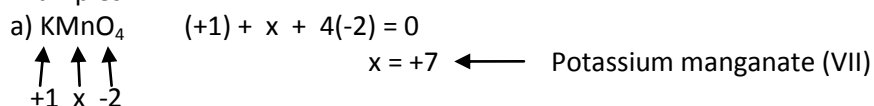
Rules for assignment of oxidation number

1. The oxidation number of an atom in its elemental state is zero. For example, the oxidation number of each atom in Mg, Cu, Na, H<sub>2</sub>, O<sub>2</sub>, P<sub>4</sub>, Fe, Zn, Br<sub>2</sub>, N<sub>2</sub>
2. The oxidation number of a monoatomic ion is equal to its charge

Ion	Na <sup>+</sup>	Mg <sup>2+</sup>	Cl <sup>-</sup>	N <sup>3-</sup>
Oxidation Number	+1	+2	-1	-3

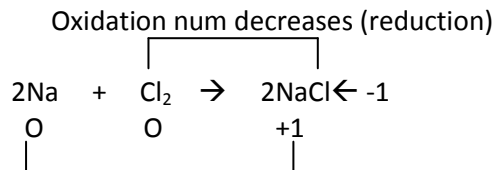
3. The oxidation number of hydrogen in a compound is always +1 except metal hydrides, where it is -1  
For example, the oxidation number of H in H<sub>2</sub>O and NH<sub>3</sub> is +1. The oxidation number of H in sodium hydride, NaH is -1
4. The oxidation number of oxygen in a compound is always -2 except in peroxide, H<sub>2</sub>O<sub>2</sub>.  
For example, the oxidation number of O in H<sub>2</sub>O and MgO is -2. However the oxidation of O in hydrogen peroxide, H<sub>2</sub>O<sub>2</sub> is -1
5. The oxidation number of fluorine in all its compound is -1  
The oxidation number of other halogen (Cl, Br, I) in their compound is -1 except when they combine with more Electronegativity elements such as oxygen and nitrogen
6. The sum of oxidation number of all the elements in the formula of a compound must be zero  
The sum of the oxidation number of all the elements in the formula of a polyatomic ion must be equal to the charge of the ion

#### • Examples



### Oxidation number in oxidation and reduction

- Example



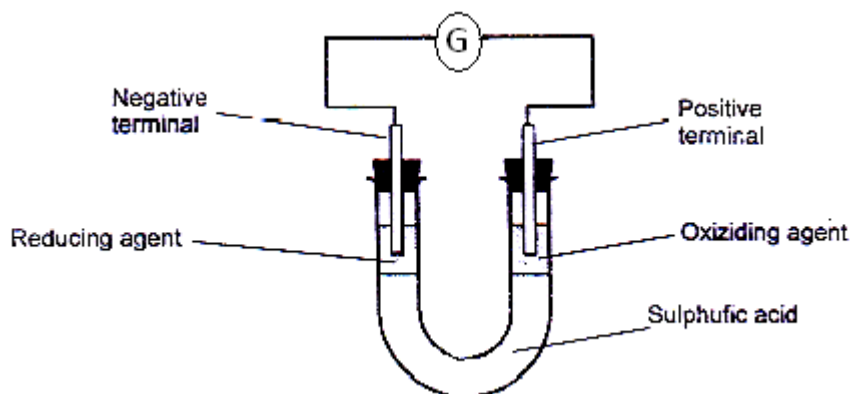
- \* Oxidized substance : Sodium
- \* Reduction substance : Chloride
- \* Oxidizing agent : Chloride
- \* Reducing agent : Sodium

### Transfer of electron at a distance

- The reducing agent and oxidizing agent are separated by an electrolyte in a U-tube
- A redox reaction takes places at the electrodes whereby the electrons are transferred through the connecting wires from the negative terminal to the positive terminal
- The electrode at which electrons are released by the reducing agent is called the negative terminal (anode)
- The electrode at which electrons are accepted by oxidizing agent is called the positive terminal (cathode)
- Examples of oxidant and reductant and the half-equation involved :

Reducing agent // releases electrons // electrodes called terminal negative (anode- undergoes oxidation)	Oxidizing agent // accept electrons // electrode is called terminal positive (cathode – undergoes reduction)
KCl : $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$	Chlorine water : $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$
KBr :	Bromine water :
KI :	Iodine water :
Mg :	$\text{KMnO}_4/\text{H}^+$ :
Zn :	$\text{K}_2\text{Cr}_2\text{O}_7/\text{H}^+$ :
Al :	$\text{CuSO}_4$ :
$\text{Sn}^{2+}$ :	$\text{Sn}^{4+}$ :

- Setup of apparatus

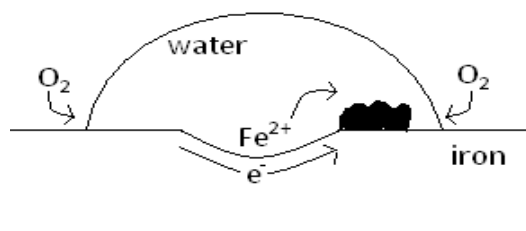


### Corrosion of metal

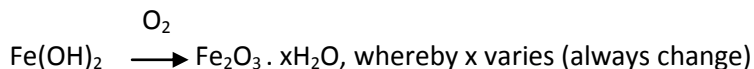
- Corrosion of metal is a redox reaction in which a metal is oxidised naturally to its ions.  
$$M \rightarrow M^{n+} + ne^{-}$$
- The more electropositive a metal is, the easier for it to corrode
- Aluminium corrodes quickly in air to form a coating of a tightly packed and non-porous aluminium oxide that can prevent aluminium from further corrosion.
- Lead and zinc are also resistant to corrosion
- Iron corrodes slowly, the coating of iron(III) oxide formed is porous and not tightly packed, cannot prevent the iron from corrosion.

### **3.2 ANALYSING RUSTING AS A REDOX REACTION**

- Rusting of iron ( $Fe \rightarrow Fe^{2+} + 2e^{-}$ ) occurs in the presence of water and oxygen



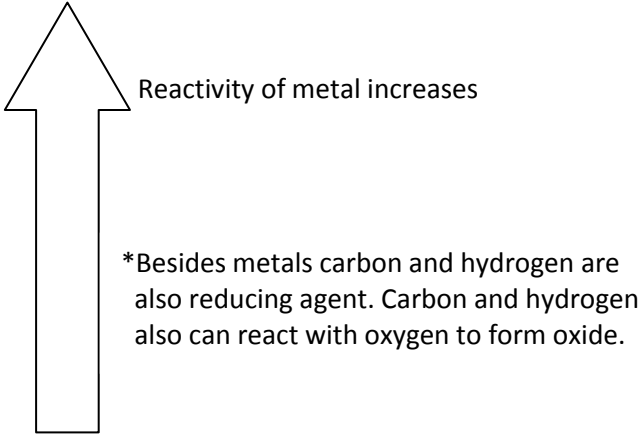
- The surface of iron in the middle of the water droplet : iron is oxidized to Iron (II) ions by loss of electrons  
Oxidation half equation : \_\_\_\_\_
- The edges of the water droplet : oxygen is reduced to hydroxide ions by gain of electrons  
Reduction half equation : \_\_\_\_\_
- The iron (II) ions combine with the hydroxide ions to form iron(II) hydroxide
- The chemical equation : \_\_\_\_\_
- The iron(II) hydroxide is then rapidly oxidized by oxygen to brown hydrated iron(III) oxide



- Rusting occurs faster in the presence of acids and salt solution which acts as electrolyte
- Way to control rusting :-
  - Alloying
  - Sacrificial protection
  - Galvanizing
  - Tin plating
  - Paints, plastics enamels and grease

### 3.3 UNDERSTANDING THE REACTIVITY SERIES OF METALS AND ITS

#### APPLICATION

- K  
Na  
Ca  
Al  
C  
Zn  
H<sub>2</sub>  
Fe  
Sn  
Pb  
Cu  
Hg  
Ag  
Au
- 
- Reactivity of metal increases
- \*Besides metals carbon and hydrogen are also reducing agent. Carbon and hydrogen also can react with oxygen to form oxide.
- Application of the reactivity series
    - a) To predict reaction involving metals  
If X metal is more reactive than metal Y, then metal X can remove oxygen from Y oxide
    - b) The extraction of metals (iron and tin)

### 3.4 ANALYSING REDOX REACTIONS IN ELECTROLYTIC AND CHEMICAL CELLS

- Electrolyte  
Anode → oxidation because releases electrons  
Cathode → reduction because accepts electrons  
EXAMPLE : Electrolysis of CuSO<sub>4</sub> solution  
Ions that presents :  $\text{CuSO}_4 \rightarrow \text{Cu}^{2+} + \text{SO}_4^{2-}$   
 $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$   
Anode :  $4\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$   
- Oxidation because release electrons  
Cathode :  $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$   
- Reduction because accepts electrons
- Chemical cell  
Positive Terminal → reduction because accepts electrons  
Negative terminal → oxidation because releases electrons.  
EXAMPLE : Chemical Cell : Mg/MgSO<sub>4</sub>//CuSO<sub>4</sub>/Cu  
Negative terminal :  $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$   
- Oxidation because release electrons  
Positive terminal :  $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$   
- Reduction because accepts electrons

