

Chemistry 2.7

Redox Chemistry

Introduction

Redox

- **Reduction – Oxidation**



LEO- Loss Electrons Oxidation

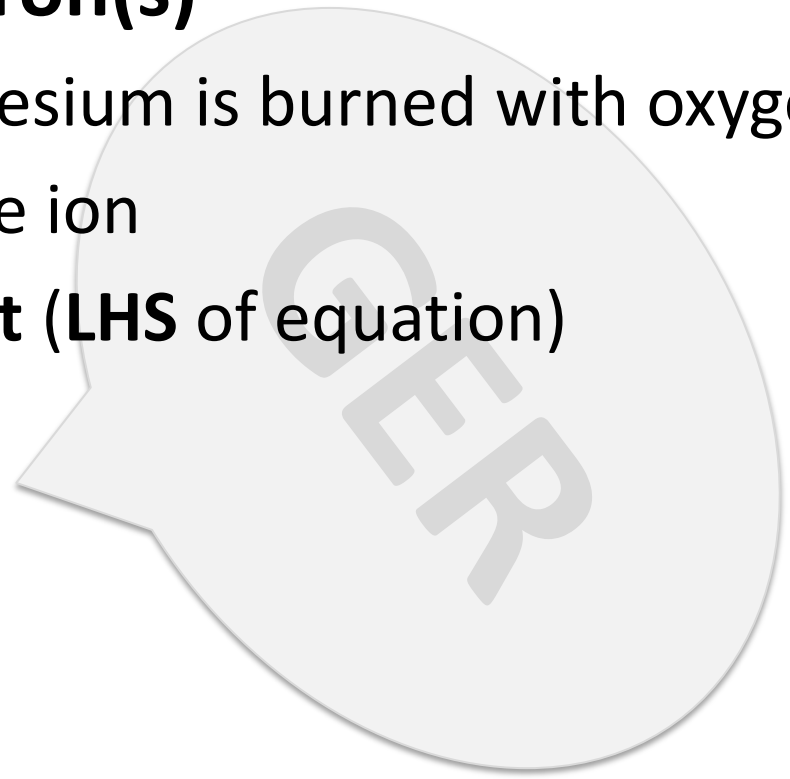
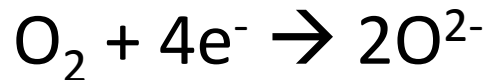
- **Oxidation** is the half of a reaction where a substance **loses electron(s)**
 - Example: When sodium metal is put into water
 - Sodium metal becomes sodium ion
 - Electron as a **product (RHS of equation)**



LEO

GER- Gain Electrons Reduction

- **Reduction** is the half of a reaction where a substance **gains electron(s)**
 - Example: When magnesium is burned with oxygen
 - Oxygen becomes oxide ion
 - Electrons as a **reactant (LHS)** of equation)



Oxidation number (ON)

- Also known as oxidation state
- It is a description of how much a substance is being oxidized
- +/- #

Example

Zn^{2+} has an oxidation number of +2

Rules	Example
When atoms exist as elements , they have an oxidation number of zero	Na, Cl ₂ , diamond, H ₂ , S ₈ ... all have an oxidation number of zero
The oxidation number of a monoatomic ion (ion with a single atom) is the same as the charge of the ion	ON of Zn ²⁺ = +2 ON of Cl ⁻ = -1 ON of O ²⁻ = -2
Hydrogen in compounds has an oxidation number of +1 (with exception of metal hydrides which is -1)	The ON of Hydrogen in H ₂ O, CH ₄ , NH ₃ are all +1
Oxygen in compounds has an oxidation number of -2 (with exception of peroxide which is -1)	The ON of oxygen in H ₂ O, MgO, CO ₂ , OCl ⁻ , are all -2
For polyatomic ions (ions containing more than one atom) the sum of oxidation number equals the charge of the ion	CO ₃ ²⁻ Since oxygen is always -2 which means carbon has to be +4 in order to form a 2- anion
The sum of the oxidation numbers of compounds is zero	CH ₄ Since hydrogen is always +1 which means carbon has to be -4 in order to have a sum of zero

Change in Oxidation number

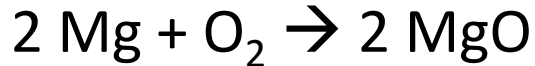
- If a substance's oxidation number increases, the substance is being oxidized. Oxidation has occurred.
- If a substance decreases in oxidation number, the substance is being reduced. Reduction has occurred.

Reaction pair

- Each redox reaction is a pair of half equations
- **Reductant** (reducing agent) = **electron provider**
 - Under goes **oxidation**
 - **Oxidation number increases**
- **Oxidant** (oxidizing agent) = **electron acceptor**
 - Under goes **reduction**
 - **Oxidation number decreases**

Example

Burning magnesium



1. Mg is a metallic substance
 1. It is in its elemental form
 2. ON = 0
2. O₂ is a molecular substance
 1. It is in its elemental form
 2. ON = 0
3. MgO is an ionic substance
 1. Consist of Mg²⁺ and O²⁻ ions
 2. The ON for Mg²⁺ is +2
 3. The ON for O²⁻ is -2

Mg's ON from
zero to +2

Mg is being **oxidized**
Mg is the **reductant**

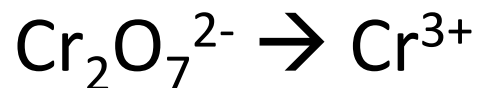
O's ON from
zero to -2

O is being **reduced**
O₂ is the **oxidant**

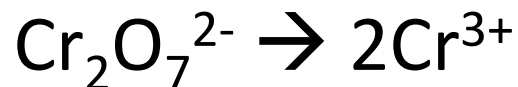
Writing $\frac{1}{2}$ equations in acidic condition

- These rules applied for acidic aqueous system
 1. Balance the atoms that are not O or H
 2. Balance the oxygen by adding H_2O
 3. Balance the hydrogen by adding H^+
 4. Balance the charge by adding electrons

Example #1



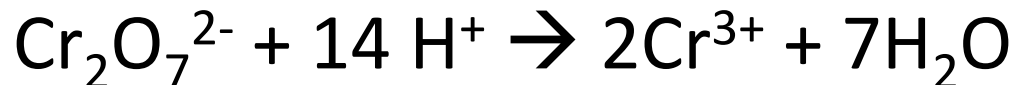
1. Balance the atoms that are not O or H



2. Balance the oxygen by adding H_2O



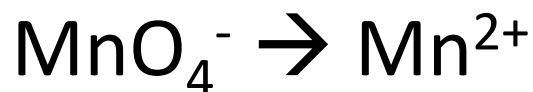
3. Balance the hydrogen by adding H^+



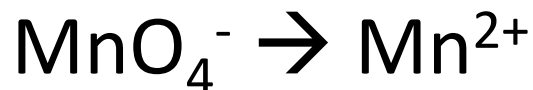
4. Balance the charge by adding electrons



Example #2



1. Balance the atoms that are not O or H



2. Balance the oxygen by adding H_2O



3. Balance the hydrogen by adding H^+



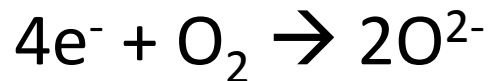
4. Balance the charge by adding electrons



Writing overall equation

- In other words, joining two half equations by getting rid of electrons.
- Take the following as an example

Aluminium being oxidized by oxygen



The **lowest common multiple** of number of electrons is 12

