Chemistry 2.7
Redox Chemistry

Introduction

Redox

- **Red**uction – **Ox**idation

LEO- **L**oss **E**lectrons **O**xidation

- **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)

$Na \rightarrow Na^+ + e^-$ Oxidation

GER- **G**ain **E**lectrons **R**eduction

- **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)

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

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_2O, 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_2O, 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

$2\ Mg + O_2 \rightarrow 2\ MgO$

1. Mg is a metallic substance
   1. It is in its elemental form
   2. ON = 0
2. O_2 is a molecular substance
   1. It is in its elemental form
   2. ON = 0
3. MgO is an ionic substance
   1. Consist of Mg^{2+} and O^{2-} ions
   2. The ON for Mg^{2+} is +2
   3. The ON for O^{2-} is -2

Mg's ON from

zero to +2

Mg is being **oxidizied**

Mg is the **reductant**

O's ON from

zero to -2

O is being **reduced**

O_2 is the **oxidant**

Writing ½ 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

$$Cr_2O_7^{2-} \rightarrow Cr^{3+}$$

1. Balance the atoms that are not O or H

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$$

2. Balance the oxygen by adding H_2O

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$$

3. Balance the hydrogen by adding H^+

$$Cr_2O_7^{2-} + 14\ H^+ \rightarrow 2Cr^{3+} + 7H_2O$$

4. Balance the charge by adding electrons

$$Cr_2O_7^{2-} + 14\ H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

Example #2

$$MnO_4^- \rightarrow Mn^{2+}$$

1. Balance the atoms that are not O or H

$$MnO_4^- \rightarrow Mn^{2+}$$

2. Balance the oxygen by adding H_2O

$$MnO_4^- \rightarrow Mn^{2+} + \mathbf{4H_2O}$$

3. Balance the hydrogen by adding H^+

$$MnO_4^- + \mathbf{8H^+} \rightarrow Mn^{2+} + 4H_2O$$

4. Balance the charge by adding electrons

$$MnO_4^- + 8H^+ + \mathbf{5e^-} \rightarrow Mn^{2+} + 4H_2O$$

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

$$Al \rightarrow Al^{3+} + 3e^-$$

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

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

$$(Al \rightarrow Al^{3+} + 3e^-) \times 4 = 4Al \rightarrow 4Al^{3+} + \cancel{12e^-}$$

$$(4e^- + O_2 \rightarrow 2O^{2-}) \times 3 = \cancel{12e^-} + 3\,O_2 \rightarrow 6O^{2-}$$

$$4Al + 3O_2 \rightarrow 4Al^{3+} + 6O^{2-}$$