REDOX

VISUAL CHEM CARDS

Nigel Freestone

Oxidation & Reduction

Definitions

Oxidation

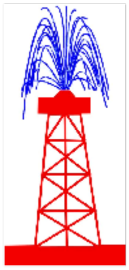
- gain in oxygen
- loss of hydrogen
- loss of electrons
- increase in oxidation number

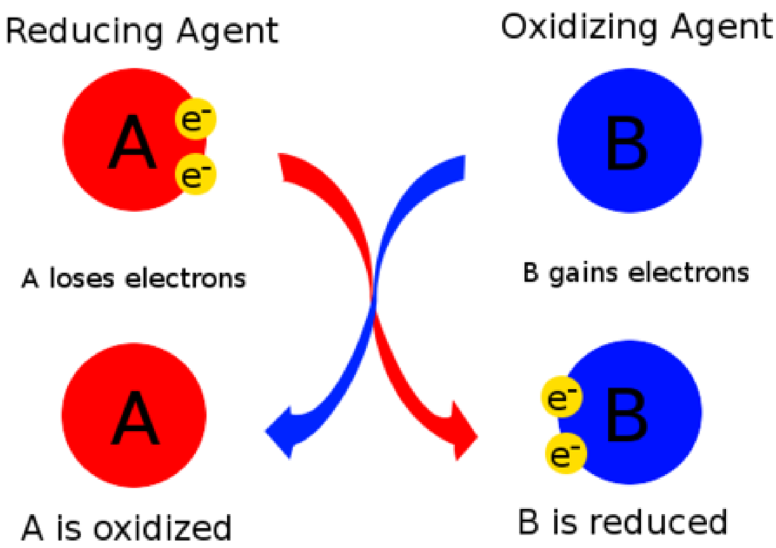
Reduction

- loss of oxygen
- gain of hydrogen
- gain of electrons
- decrease in oxidation number

OIL
Oxidation **I**s **L**oss

RIG
Reduction **I**s **G**ain





A reducing agent **reduces** other substances and **loses** electrons; therefore, its oxidation state **increases**.

An oxidizing agent **oxidizes** other substances and **gains** electrons; therefore, its oxidation state **decreases**

Oxidation & Reduction

An oxidising agent:

- is normally a non-metal or positive ion;
- cause oxidation reactions to take place;
- gains electrons from other atoms or ions (is itself reduced).

For example, chlorine and copper ions are both oxidising agent which are themselves reduced as follows:

$Cl_{2(g)} + 2e^- \longrightarrow 2Cl^-_{(aq)}$

$Cu^{2+}_{(aq)} + 2e- \longrightarrow Cu_{(s)}$

The strongest oxidising agents are highly electronegative elements like the halogens (Group 7).

Oxidising agents are frequently used because of the effectiveness with which they can kill fungi and bacteria, and can inactivate viruses.

Dichromate ($Cr_2O_7^{2-}$) and permanganate (MnO_4^-) ions are strong oxidising agents in acidic solutions.

Hydrogen peroxide (H_2O_2) is an example of a molecule which is a strong oxidising agent.

A reducing agent:

- is usually a metal or a negative ion;
- loses (donates) electrons to another element or ion (reducing the other species);
- is itself oxidised;

For example, sodium is a reducing agent which is itself oxidised as follows:

$Na_{(s)} \longrightarrow Na^+_{(aq)} + e^-$

The strongest reducing agents are the alkali metals (Group 1) as they have low electronegativities and lose electrons very easily.

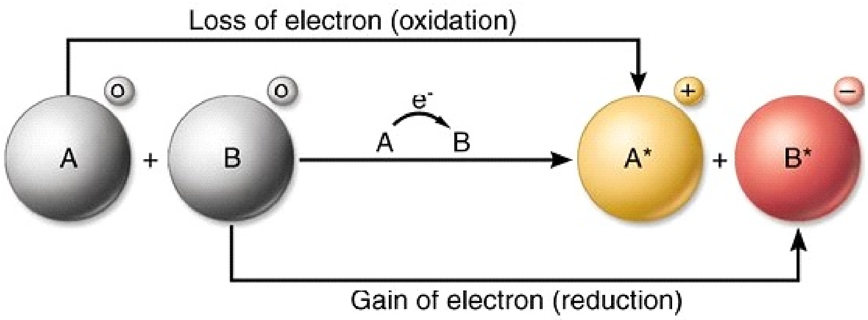
Some molecules such as carbon monoxide (CO) are also used in the chemical industry as reducing agents to help extract metals.

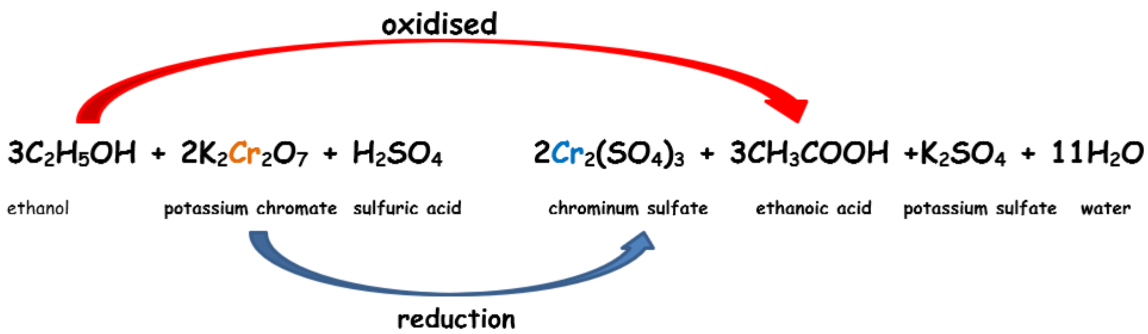
Redox

Oxidation-reduction (**redox**) reaction is a type of chemical reaction that involves a transfer of electrons between two species.

An oxidation-reduction reaction is any chemical reaction in which the oxidation number of a molecule, atom, or ion changes by gaining or losing an electron.

Redox reactions are common and vital to some of the basic functions of life, including photosynthesis, respiration, combustion, and corrosion or rusting.





Potassium chlorate is the **oxidising agent**

Oxidation numbers : Cr in CrO_7^{2-} = +6

: Cr in $Cr_2(SO_4)_3$ = +3

Chromium is **reduced**

Ethanol (C_2H_5OH) is oxidised (addition of oxygen) to ethanoic acid (CH_3COOH).

Oxidation Numbers

Oxidation numbers which can be **positive, negative or zero** in value, are assigned to elements and simply represent the number of electrons lost or gained by an atom of that element in the compound.

Oxidation numbers are a 'bookkeeping tool' to allow chemists to keep track of electrons in reactions, identify species that are being oxidised/reduced and to balance redox reactions.

Assigning Oxidation Numbers

Although difficult to define, oxidation numbers are surprisingly easy to assign, using the following rules:

Rule 1: uncombined elements - for example, $Fe_{(s)}$, $O_{2\ (g)}$ have an oxidation number of zero.

Rule 2: the oxidation state of a monatomic ion is the same as the charge on the ion. Thus Ni^{2+} has an oxidation number of +2 and O^{2-} has a oxidation number of -2. Note the charge is given first then the number, whereas the number, then the charge is the used for ions.

Rule 3: the sum of the oxidation numbers of the elements in a neutral compound is zero, whilst equals the charge on a polyatomic ion equals the charge on the ion.

Rule 4: The oxidation number of group 1 and group 2 metals in compounds is +1 and +2, respectively. Fluorine in compounds always has an oxidation number of -1, whilst the remaining halogen elements have an oxidation number of -1 unless bonded to oxygen or fluorine.

Rule 5: Oxygen in compounds usually has an oxidation number of -2. However, in peroxides (O^{2-}) oxygen has an oxidation number of -1 and in combination with fluorine, oxygen has an oxidation number of +1.

Determining Oxidation Numbers

Determining Oxidation Numbers

Oxidation numbers (element underlined) can be readily determined by following the rules outlined above.

CuO

oxidation number of Cu + oxidation number of O = 0 (Rule 3)

oxidation number of Cu = 0- oxidation number of O

oxidation number of Cu = 0 - -2 (Rule 5)

= **+2**

$Cr_2O_7^{2-}$

2 x oxidation number of Cr + (7 x oxidation number of O) = -2 (Rule 2)

2 x oxidation number of Cr = -2- (7 x oxidation number of O)

2 x oxidation number of Cr = - 2 – {7 x -2 (Rule 2)} = +12

Oxidation number of Cr = +12/2 = **+6**

$MgSF_6$

oxidation number of Mg + oxidation number of S + (6 x oxidation number of F) = 0 (Rule 3)

oxidation number of S = 0 – oxidation number of Mg – (6 x oxidation number of F)

oxidation number of S = 0 – +2 (Rule 4) – {6 x -1 (Rule 4)

= 0 – 2 +6

= **+4**