Redox Reactions

Combustion, explosions, rusting, rotting, breathing





Oxidation numbers

- The **oxidation state** or **oxidation number**, is an indicator of the degree of oxidation (loss of electrons) of an atom in a chemical compound.
- The oxidation state, which may be positive, negative or zero, is the *hypothetical* charge that an atom would have if all bonds to atoms of different elements were 100% ionic, with no covalent component. This is never exactly true for real bonds.

Different oxidation states of the same element in compounds are the reason for different properties.

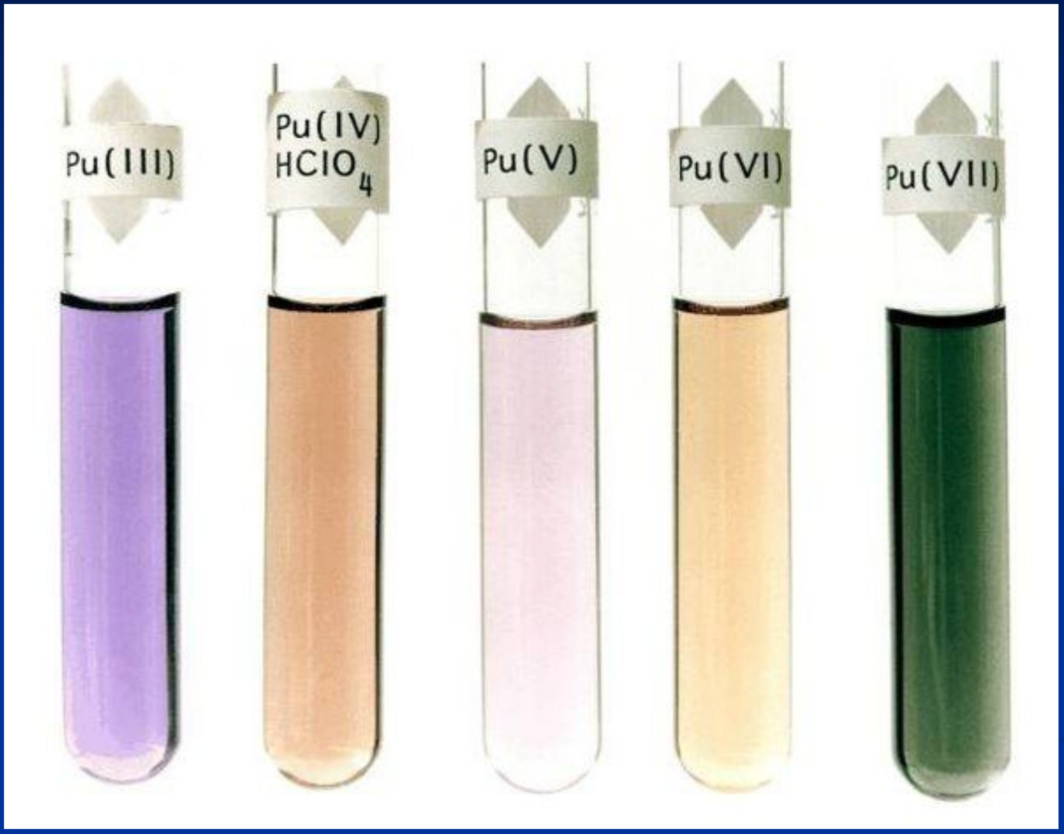
$Fe(OH)_2$

iron is in oxidation state +2

$Fe(OH)_3$

iron is in oxidation state +3

Plutonium oxidation states



Determining the oxidation state or number

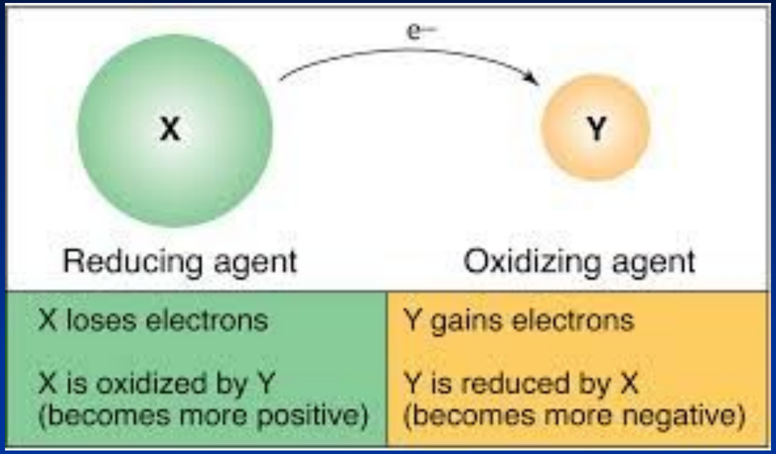
- Any pure element—even if it forms diatomic molecules like chlorine (Cl_2)—has an oxidation state of zero.
- For monatomic ions, the oxidation state is the same as the charge of the ion. E.g., the chloride anion (Cl^-) has an oxidation state of −1, whereas the lithium cation (Li^+) has an oxidation state of +1.
- The sum of oxidation states for all atoms in a polyatomic ion is equal to the charge of the ion. Thus, the oxidation state of one element can be calculated from the oxidation states of the other elements.
- The sum of the oxidation states of all atoms in a neutral molecule must be zero.

Rules based on electronegativity

- Fluorine in compounds has an oxidation state of −1.
- Halogens other than fluorine have an oxidation state of −1 except when they are bonded to oxygen, to nitrogen, or to another halogen that is more electronegative.
- Hydrogen has an oxidation state of +1 except when bonded to more electropositive elements such as sodium, aluminium, and boron.
- In compounds, oxygen typically has an oxidation state of −2.
- Alkali metals have an oxidation state of +1 in virtually all of their compounds.
- Alkaline earth metals have an oxidation state of +2 in virtually all of their compounds.

Definitions:

- **Oxidation** is *the loss of electrons*.
- **Reduction** is *the gain of electrons.*



- Any reaction involving the transfer of electrons is an **ox**idation-**red**uction (or **red**ox) reaction
- Oxidation cannot take place without reduction.
- During a redox reaction, the oxidation numbers of reactants will change.

Oxidizing agents

- Substances that have the ability to oxidize other substances (cause them to lose electrons) are known as oxidizing agents, oxidants, or oxidizers. The oxidizing agent removes electrons from another substance, and is thus itself reduced. And, because it "accepts" electrons, the oxidizing agent is also called an electron acceptor.
- Oxygen is the quintessential oxidizer.
- Substances with elements in high oxidation states (e.g., concentrated sulfuric acid H_2SO_4, potassium permanganate $KMnO_4$, potassium dichromate(VI) $K_2Cr_2O_7$, manganese(IV) oxide MnO_2)

or else highly electronegative elements (O_2, F_2, Cl_2, Br_2) that can gain extra electrons by oxidizing another substance.

In the Thermit reaction, shown here, which substance is reduced and which is oxidized?

Aluminium (Al) removes the oxygen atoms from the iron(III) oxide (Fe_2O_3). The heat needed to start the reaction is usually provided by a magnesium fuse.

$2Al + Fe_2O_3 \quad Al_2O_3 + 2Fe$

The iron(III) ions are reduced and the aluminium ions are oxidized.

This reaction produces much heat. It is used in incendiary weapons and in underwater welding.

Reducing agents

- Substances that have the ability to reduce other substances (cause them to gain electrons) are said to be reductive or reducing and are known as reducing agents, reductants, or reducers. They transfer electrons to another substance, and are thus oxidized. And, because it "donates" electrons, the reducing agent is also called an electron donor.
- Electropositive elemental metals, such as Li, Na, Mg, Fe, Zn, and Al. These metals donate or *give away* electrons readily.
- C, CO, H_2

For any equation to be balanced:

1. The number of atoms of each type on the left side of the arrow must equal the number of atoms of each type to the right of the arrow.

2. The total charges on the left side of the arrow must equal the total charges to the right of the arrow.

3. The electrons lost (during oxidation) must equal the electrons gained (during reduction).

Types of redox reactions

- 1. Combination

General equation: $A+B \rightarrow AB$

$2\,H_2 + O_2 = 2\,H_2O$

- 2. Decomposition

General equation: $AB \rightarrow A + B$

$2Ag_2O = 4Ag + O_2$ (heating)

$(NH_4)_2Cr_2O_7 = N_2 + Cr_2O_3 + 4H_2O$ (heating)

Decomposition reactions are the reverse of combination reactions, meaning they are the breakdown of a chemical compound into its component elements.

Balancing the equations

$K^+Mn^{+7}O_4^{-2} + K^+Cl^- + H_2SO_4 = Cl_2^0 + Mn^{+2}SO_4 + K_2SO_4 + H_2O$

- $Mn^{+7} + 5e- = Mn^{+2}$ reduction 2
- $2Cl^- - 2e- = Cl_2^0$ oxidation 10 5

$2\,KMnO_4 + 10KCl + 8H_2SO_4 = 5Cl_2 + 2MnSO_4 + 6K_2SO_4 + 8H_2O$

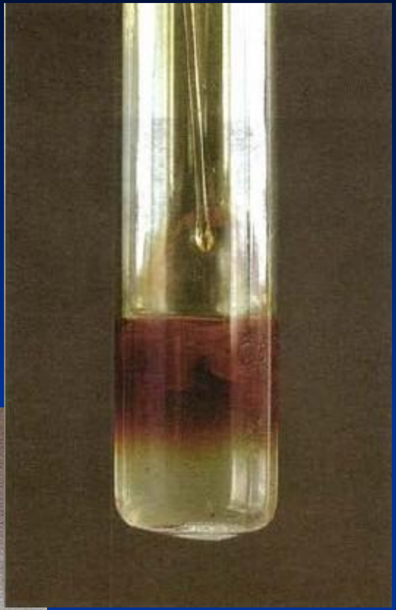
Types of redox reactions

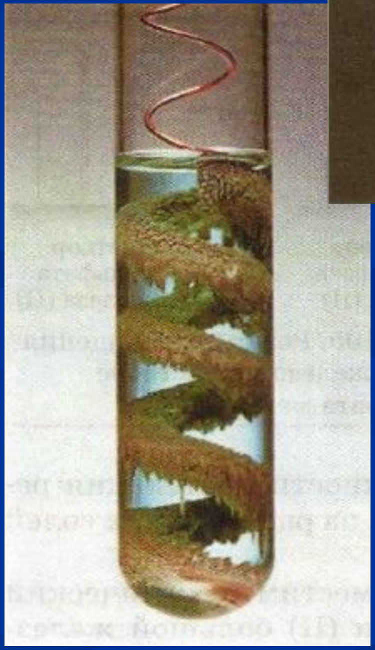
- **Single Displacement**

General equation: A + BC → AB + CA

$Cl_2 + 2\,NaBr = 2\,NaCl + Br_2$

$2AgNO_3 + Cu = Cu(NO_3)_2 + 2Ag$





Types of redox reactions

- **Combustion**

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + Q$

Combustion reactions always involve oxygen and an organic fuel.

Types of redox reactions

- **Disproportionation**
- **The same substances are both oxidized and reduced. These are known as disproportionation reactions.**

$2\ H_2O_2(aq) \rightarrow 2\ H_2O(l) + O_2(g)$

In H_2O_2, oxygen has an oxidation state of -1.

In H_2O, its oxidation state is -2,

and it has been reduced.

In O_2 its oxidation state is 0,

and it has been oxidized.

Oxygen has been both oxidized and reduced in the reaction, making this a disproportionation reaction.

- Chemistry IB – Chapter 10, pp.302-335

- Balance the equations

$NH_3 + O_2 = NO + H_2O$

$CuO + NH_3 = Cu + N_2 + H_2O$

$NO_2 + H_2O = HNO_3 + NO$

$NH_4NO_3 = N_2O + H_2O$

$K_2Cr_2O_7 + KI + H_2SO_4 = Cr_2(SO_4)_3 + K_2SO_4 + I_2 + H_2O$

$KMnO_4 + HCl = Cl_2\uparrow + MnCl_2 + KCl + H_2O$