Inorganic Chemistry Department
General and Inorganic Chemistry

Oxidation-Reduction Reaction

In specialty 226 Pharmacy, industry pharmacy

Ye.Ya.Levitin,
Doctor of Pharmaceutical Sciences, Professor

Plan

1. **The concept of oxidation and reduction reaction.**

2. **Redox reaction equation balancing methods.**

3. **Electron-ion (half reaction) method**.

4. **Types of redox reactions**.

5. **Factors affecting the course of redox reactions.**

6. **Measurement of Oxidation and Reduction Tendencies.**

7. **The Activity Series of Metals.**

Actuality of theme

The world around us is a giant chemical laboratory in which thousands of reactions take place every second, mostly redox, and as long as they exist, these reactions, as long as there are conditions for their flow, perhaps all the glory surrounding us, life itself is possible.

Aim: To expand and consolidate knowledge of the degree of oxidation, oxidizing agents and reducing agents; form an ability to equalize reactions using the electronic balance method; to acquaint with the process of breathing, the meaning of breathing, the process of gas exchange in the lungs and tissues; improve the ability of students to generalize, draw conclusions, analyze, compare, establish a causal relationship between the composition, structure and properties of substances; creation of conditions for awareness of the need for a competent attitude to their health and the environment.

Specific goals:
- The concept of oxidizing and reduction process.
- To be able to select stoichiometric coefficients in redox reactions.
- Be able to determine the possibility of redox reactions.

Photosynthesis.

Combustion.

Breath.

Rotting.

Corrosion.

Pyrotechnics.

Bibliographical Guidance

General and inorganic chemistry: Textbook for students of higher schools Ye.Ya.Levitin, I.O.Vedernikova.– Kharkiv:Publishing House of NUPH :Golden Pages, 2009. – 360 p.

Laboratory Practicum in General and inorganic chemistry.I module/ Ye.Ya.Levitin, I.O.Vedernikova et al.– Kharkiv,2015. – 79 p.

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General and inorganic chemistry:The lecture courses for the students of pharmaceutical universities. I Module / Ye.Ya.Levitin, I.O.Vedernikova et al.– Kharkiv: NUPH Publishers, 2014. – 87 p.

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The collection of theoretical questions and tests in inorganic chemistry:for the students of higher school. / Ye.Ya.Levitin, I.O.Vedernikova.– Kharkiv: NUPH Publishers, 2007. – 76 p.

Redox (*reduction-oxidation*) reactions include all chemical reactions in which atoms have their oxidation state changed. This can be either a simple redox process, such as the oxidation of carbon to yield carbon dioxide (CO_2), or complex process such as the oxidation of glucose in the human body through a series of complex electron transfer processes.

The oxidation state is the hypothetical charge that an atom would have if we assume that the molecule is composed of ions and the charge of the molecule is 0.

$$H_2S$$
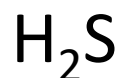
S^{-2} is the oxidation state of S

S^{2-} is the charge of sulphide ion

1. Alkaline, alkaline earth metals, Zn, Al have constant oxidation state, it is equal to the group number.

2. Simple substances (S, Fe, O_2) have oxidation state "0".

3. The basic oxidation state of Hydrogen is **+1.** Its oxidation state is **-1** only in hydrides of alkaline and alkaline earth metals.
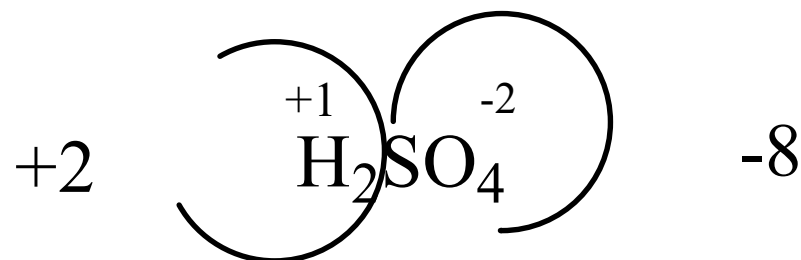
NaH, CaH_2

4. The basic oxidation state of Oxygen is -2 and -1 in peroxides only.

SO_2, H_2O_2, Ba_2O_2
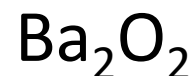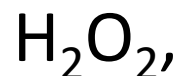
5. The sum of all oxidation states in the compound is 0.

Let's calculate the oxidation state of the elements in H_2SO_4.

The oxidation state of H is +1 (rule 3). The total of the oxidation state for two H atoms is +2.

$$+2 \qquad \overset{+1}{H_2}\overset{-2}{S O_4} \qquad -8$$

The oxidation state of O is −2 (rule 4). The total oxidation state for four atoms O is −8:

$$2(+1) + x + 4(-2) = 0$$
$$x = +6$$
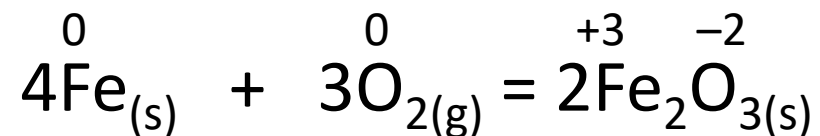
Eks is equal to +6.

The oxidation state of S is +6.

Reducing agent is a substance, which gives the electrons. Its oxidation state increases.

Oxidizing agent is a substance, which takes the electrons. Its oxidation state decreases.
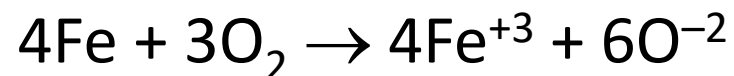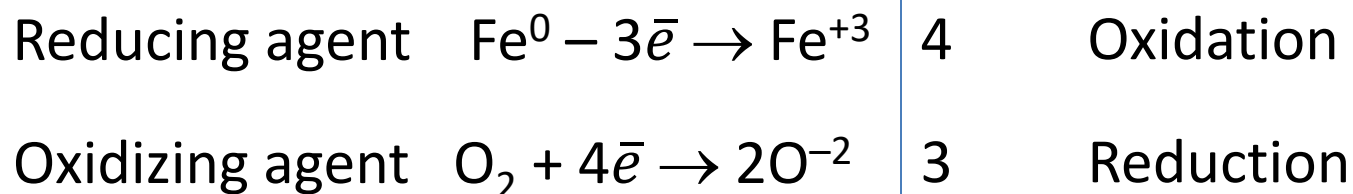
Reduction is the gain of electrons to an atom, molecule or ion, accompanied by decreasing of the oxidation state of element.

The oxidation is the loss of electrons by atom, molecule or complex ion, accompanied by increasing of the oxidation state of element.

When an iron object is exposed to the atmosphere, it rusts:

$$\overset{0}{4Fe}_{(s)} + \overset{0}{3O}_{2(g)} = \overset{+3}{2Fe_2}\overset{-2}{O}_{3(s)}$$

In this reaction iron combines with oxygen. Originally the term "oxidation" was applied to reactions in which a substance combines with oxygen:

Reducing agent	$Fe^0 - 3\bar{e} \rightarrow Fe^{+3}$	4	Oxidation
Oxidizing agent	$O_2 + 4\bar{e} \rightarrow 2O^{-2}$	3	Reduction

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

The number of electrons accepted by the oxidizing agent is equal to the number of electrons given by the reducing agent.

The oxidation state of iron before the reaction was zero. It increases from 0 to +3. The oxidation state of oxygen decreases from 0 to −2.
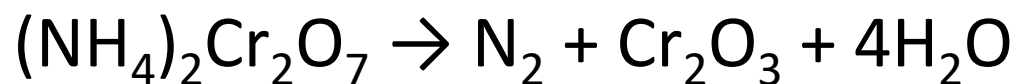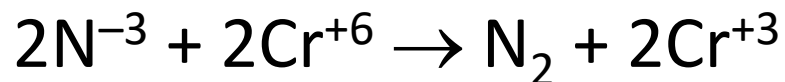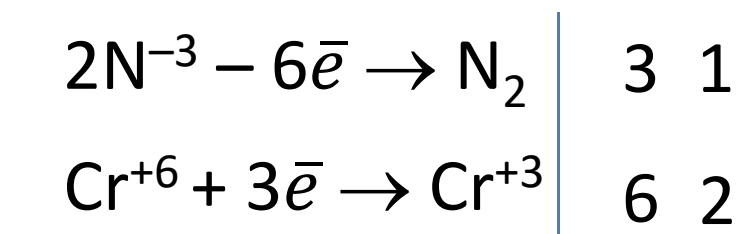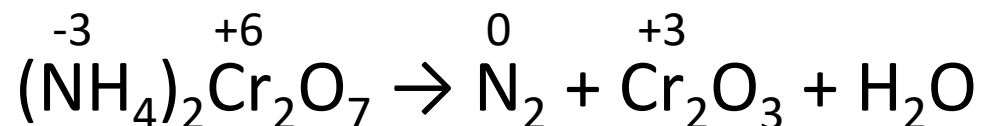
Dependence of red-ox properties on the degree of oxidation

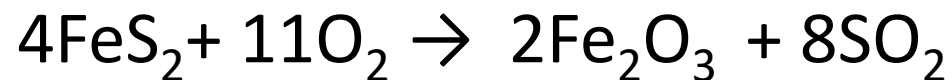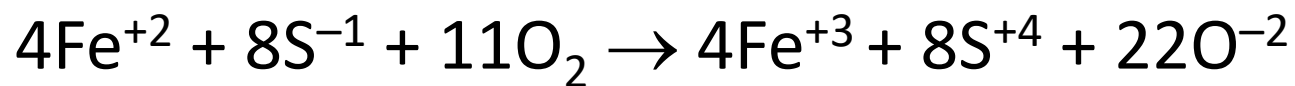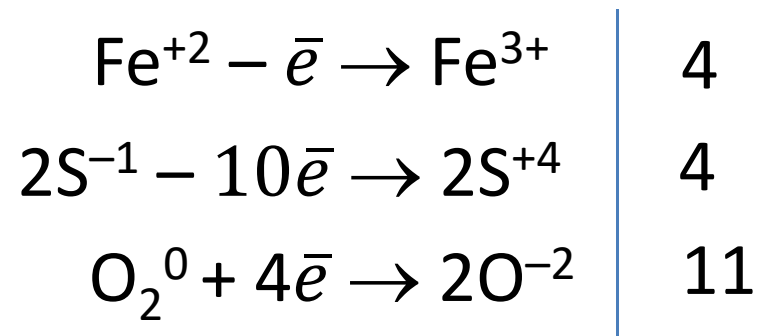-2 $\quad$ H_2S $\quad$ $3s^2 3p^6$ $\quad$ Only reduction properties

0 $\quad$ S $\quad$ $3s^2 3p^4$ $\quad$ Redox duality

$+4$ $\quad$ SO_2 $\quad$ $3s^2 3p^0$ $\quad$ Redox duality

$\quad\quad\quad$ Na_2SO_3

$+6$ $\quad$ H_2SO_4 $\quad$ $3s^0 3p^0$ $\quad$ Only oxidation properties

Redox reaction equation balancing methods

There are two methods of selection coefficients in redox reaction.

The electron balance is the universal, the only possible method for reactions of fusion, decomposition taking place in the non-aqueous medium.

$$\overset{-3}{(NH_4)_2}\overset{+6}{Cr_2}O_7 \rightarrow \overset{0}{N_2} + \overset{+3}{Cr_2}O_3 + H_2O$$

$$2N^{-3} - 6\bar{e} \rightarrow N_2 \quad | \quad 3 \quad 1$$

$$Cr^{+6} + 3\bar{e} \rightarrow Cr^{+3} \quad | \quad 6 \quad 2$$

$$2N^{-3} + 2Cr^{+6} \rightarrow N_2 + 2Cr^{+3}$$

$$(NH_4)_2Cr_2O_7 \rightarrow N_2 + Cr_2O_3 + 4H_2O$$

$$\_FeS_2 \;+\; \_O_2 \rightarrow \_Fe_2O_3 \;+\; \_SO_2$$

$$
\begin{array}{r|l}
Fe^{+2} - \bar{e} \rightarrow Fe^{3+} & 4 \\
2S^{-1} - 10\bar{e} \rightarrow 2S^{+4} & 4 \\
O_2{}^0 + 4\bar{e} \rightarrow 2O^{-2} & 11
\end{array}
$$

$$4Fe^{+2} + 8S^{-1} + 11O_2 \rightarrow 4Fe^{+3} + 8S^{+4} + 22O^{-2}$$

$$4FeS_2 + 11O_2 \rightarrow 2Fe_2O_3 + 8SO_2$$

Electron-ion (Half reaction) Method can be used for aqueous solutions only.
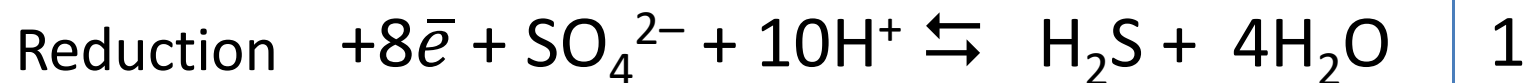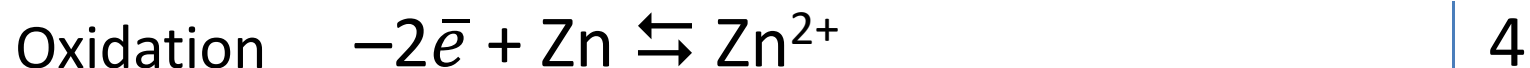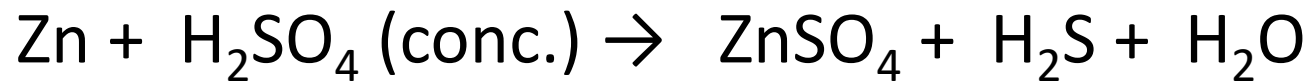
In this method:

✓ the simple substances (Cl_2, O_2, Fe, S),

✓ weak electrolytes (H_2O, H_2O_2, H_2S, HNO_2),

✓ the slightly soluble in water substances (PbO_2, MnO_2, $Fe(OH)_2$)

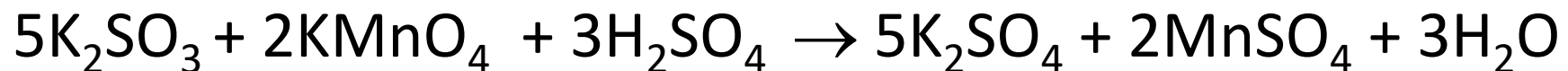✓ gases (NO_2, NO, SO_2)

are written in the molecular forms.

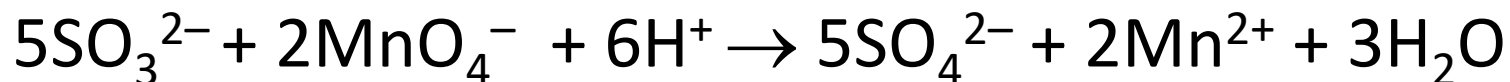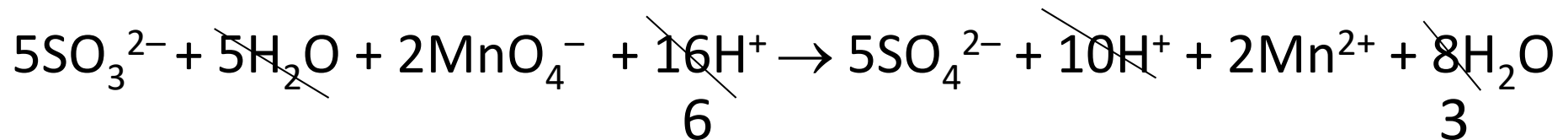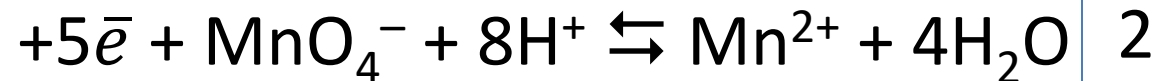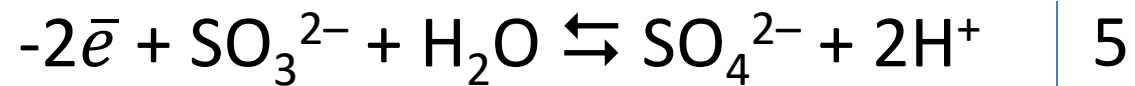Strong electrolytes are written in ionic forms (S^{2-}, SO_4^{2-}, NO_3^-, MnO_4^-, Mn^{2+}).

How to balance redox reactions in acidic medium

For an acidic solution add two H^+ for every surplus O atom. To the opposite side of the halt-equation, add one H_2O molecule.

$$Zn + H_2SO_4 \text{ (conc.)} \rightarrow ZnSO_4 + H_2S + H_2O$$

Oxidation	$-2\bar{e} + Zn \leftrightarrows Zn^{2+}$	4
Reduction	$+8\bar{e} + SO_4^{2-} + 10H^+ \leftrightarrows H_2S + 4H_2O$	1

$$4Zn + SO_4^{2-} + \boxed{10H^+} \rightarrow 4Zn^{2+} + \boxed{H_2}S + \boxed{4H_2O}$$

$$4Zn + 5H_2SO_4 \text{ (conc.)} \rightarrow 4ZnSO_4 + H_2S + 4H_2O$$

$$E_m(Red) = \frac{M_{Red}}{n\bar{e}} = \frac{65}{2} = 32.5$$

$$K_2SO_3 + KMnO_4 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + H_2O$$

$$-2\bar{e} + SO_3^{2-} + H_2O \leftrightarrows SO_4^{2-} + 2H^+ \quad \bigg| \quad 5$$

$$+5\bar{e} + MnO_4^- + 8H^+ \leftrightarrows Mn^{2+} + 4H_2O \quad \bigg| \quad 2$$

$$5SO_3^{2-} + \cancel{5H_2O} + 2MnO_4^- + \cancel{16}H^+ \rightarrow 5SO_4^{2-} + \cancel{10}H^+ + 2Mn^{2+} + \cancel{8}H_2O$$
$$\qquad\qquad\qquad\qquad 6 \qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad 3$$

$$5SO_3^{2-} + 2MnO_4^- + 6H^+ \rightarrow 5SO_4^{2-} + 2Mn^{2+} + 3H_2O$$

$$5K_2SO_3 + 2KMnO_4 + 3H_2SO_4 \rightarrow 5K_2SO_4 + 2MnSO_4 + 3H_2O$$
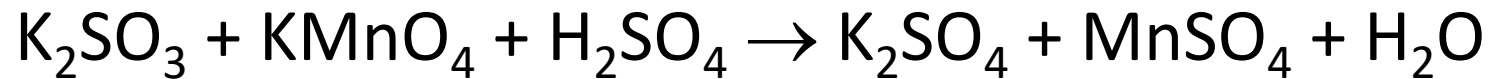
$$E_m(Ox) = \frac{M_{Ox}}{n\bar{e}} = \frac{158}{5} = 31.6$$

How to balance redox reactions in neutral and basic medium.

In neutral and basic solution add one H_2O for every surplus O atom. To the opposite side of the halt-equation add two OH^-.

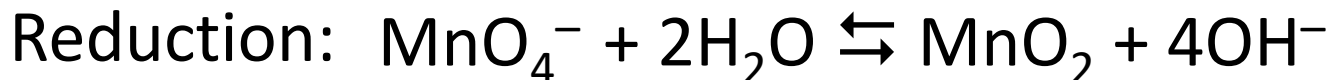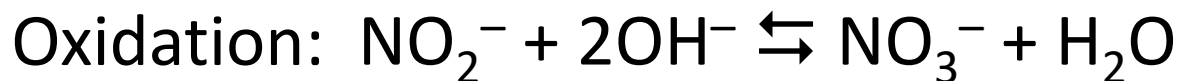$$KNO_2 + KMnO_4 + H_2O \leftrightarrows KNO_3 + MnO_2 + KOH$$
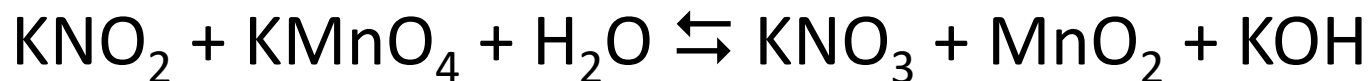
Step 1 The oxidation state of N in KNO_2 is +3 and in KNO_3 it is +5. The oxidation state of Mn decreases from +7 to +4 in the net reaction.

Oxidation: $NO_2^- \leftrightarrows NO_3^-$

Reduction: $MnO_4^- \leftrightarrows MnO_2$

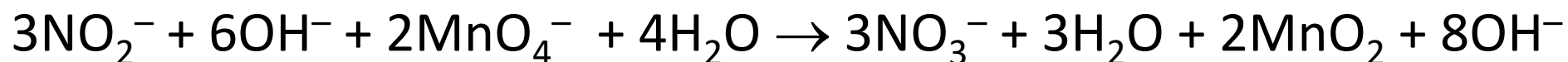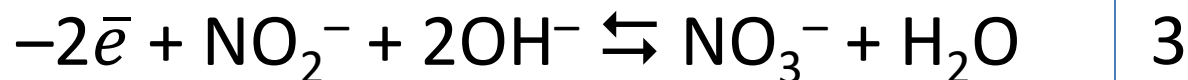**Step 2** In neutral and basic solution add one H_2O for every surplus O atom. To the other side add two OH^-.

Oxidation: $NO_2^- + 2OH^- \leftrightarrows NO_3^- + H_2O$

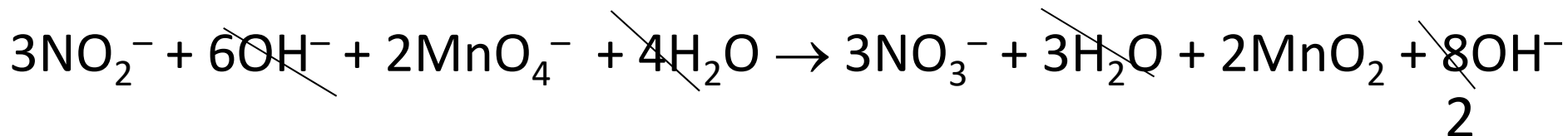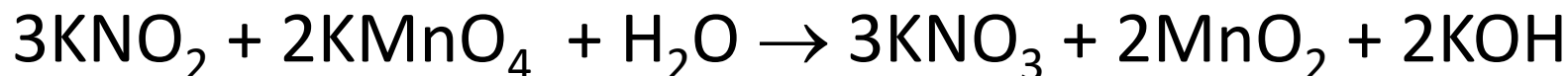Reduction: $MnO_4^- + 2H_2O \leftrightarrows MnO_2 + 4OH^-$

Step 3 In the left side of the oxidation half-equation the total number of the charges is −3, in the right side it is −1, and therefore we must take two electrons. In the left side of the next reduction half-equation the total number of the charge is −1, in the right side it is −4. Therefore we must give three electrons.
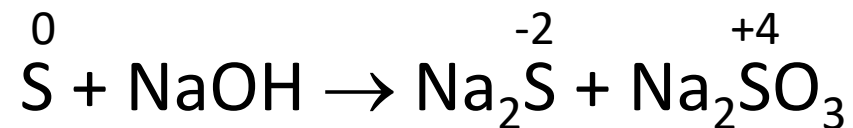
$$-2\bar{e} + NO_2^- + 2OH^- \leftrightarrows NO_3^- + H_2O \quad \Big| \quad 3$$

$$+3\bar{e} + MnO_4^- + 2H_2O \leftrightarrows MnO_2 + 4OH^- \quad \Big| \quad 2$$

$$3NO_2^- + 6OH^- + 2MnO_4^- + 4H_2O \rightarrow 3NO_3^- + 3H_2O + 2MnO_2 + 8OH^-$$

Step 4 Simplify.

$$3NO_2^- + 6OH^- + 2MnO_4^- + 4H_2O \rightarrow 3NO_3^- + 3H_2O + 2MnO_2 + 8OH^-$$
$$2$$

A summery equation of this reaction:

$$3KNO_2 + 2KMnO_4 + H_2O \rightarrow 3KNO_3 + 2MnO_2 + 2KOH$$

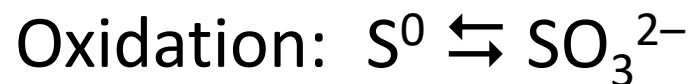$$\overset{0}{S} + NaOH \rightarrow Na_2\overset{-2}{S} + Na_2\overset{+4}{SO_3}$$

Determine the oxidation states of S before and after the reaction. The oxidation state of S in a free element is 0, in Na_2S is -2 and in Na_2SO_3 it is $+4$.
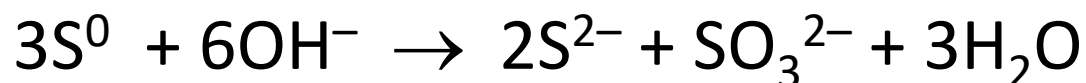
In this reaction sulphur undergo both: oxidizing agent and reducing agent, therefore this reaction called **a disproportionation reaction**.
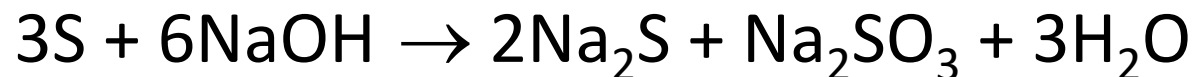
Oxidation: $S^0 \leftrightarrows SO_3^{2-}$

Reduction: $S^0 \leftrightarrows S^{2-}$

| Step 2 | Add three H_2O to the right side of the first half-equation and six OH^- to the left side. |

Oxidation: $S^0 + 6OH^- \rightleftarrows SO_3^{2-} + 3H_2O$

Reduction: $S^0 \rightleftarrows S^{2-}$
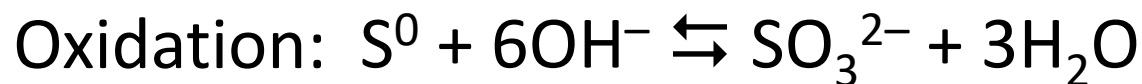
| Step 3 | Reducing agent gives fore electrons and oxidizing agent takes two electrons. |

$$-4\bar{e} + S^0 + 6OH^- \rightleftarrows SO_3^{2-} + 3H_2O \quad | \quad 1$$

$$+2\bar{e} + S^0 \rightleftarrows S^{2-} \quad | \quad 2$$

$$3S^0 + 6OH^- \rightarrow 2S^{2-} + SO_3^{2-} + 3H_2O$$

| Step 4 | A summery equation of this reaction: |

$$3S + 6NaOH \rightarrow 2Na_2S + Na_2SO_3 + 3H_2O$$

$$H_2O_2 + K_2Cr_2O_7 + H_2SO_4 = O_2 + Cr_2(SO_4)_3 + K_2SO_4 + H_2O$$

$$-2\bar{e} + H_2O_2 \rightleftarrows O_2 + 2H^+ \qquad\qquad 3$$

$$+6\bar{e} + Cr_2O_7^{2-} + 14H^+ \rightleftarrows 2Cr^{3+} + 7H_2O \quad 1$$

$$3H_2O_2 + Cr_2O_7^{2-} + 14H^+ \rightarrow 3O_2 + 6H^+ + 2Cr^{3+} + 7H_2O$$
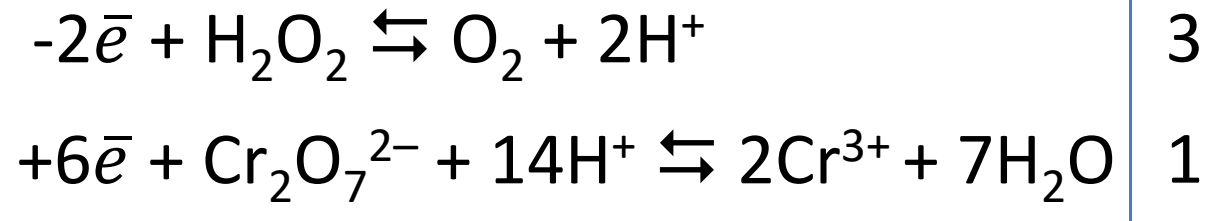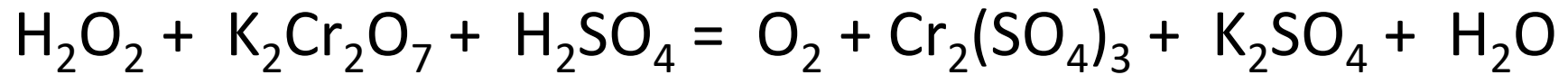
After shortening of the similar ions:

$$3H_2O_2 + Cr_2O_7^{2-} + 8H^+ \rightarrow 3O_2 + 2Cr^{3+} + 7H_2O$$

A summery equation of this reaction:

$$3H_2O_2 + K_2Cr_2O_7 + 4H_2SO_4 = 3O_2 + Cr_2(SO_4)_3 + K_2SO_4 + 7H_2O$$
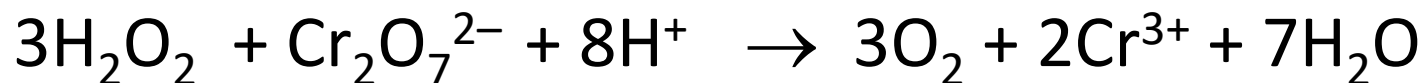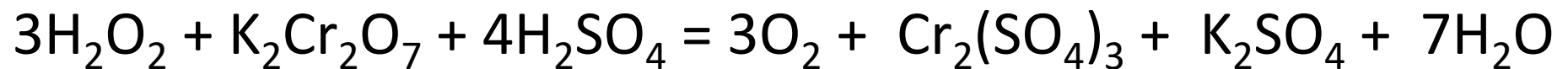
Types of redox reactions

There are three main types of oxidation-reduction reactions:

– intermolecular,

– intramolecular,

– disproportionation (dismutation).

Intermolecular reactions are called red-ox reactions, in which **elements** that change the oxidation state **are in different molecules:**

$$\overset{+3}{5NaNO_2} + \overset{+7}{2KMnO_4} + 3H_2SO_4 = \overset{+5}{5NaNO_3} + \overset{+2}{2MnSO_4} + K_2SO_4 + 3H_2O$$
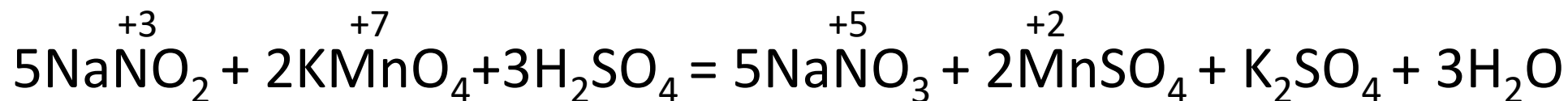
$$\overset{-2}{2H_2S} + \overset{+4}{H_2SO_3} = \overset{0}{3S} + 3H_2O$$

$$\overset{-1}{NaCl} + \overset{+1}{NaClO} + H_2SO_4 = \overset{0}{Cl_2} + Na_2SO_4 + H_2O$$
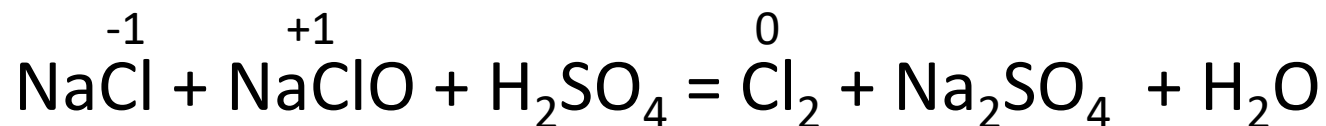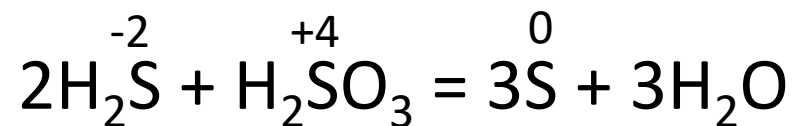
One atom of a molecule is oxidized and **other atom of same molecule is reduced** then it is **intramolecular** redox reaction.

$$\overset{-3}{(NH_4)}_2\overset{+6}{Cr_2}O_7 = \overset{0}{N_2} + \overset{+3}{Cr_2}O_3 + 4H_2O$$

$$\overset{-3}{NH_4}\overset{+3}{NO_2} = \overset{0}{N_2} + 2H_2O$$

$$\overset{-3}{NH_4}\overset{+5}{NO_3} = \overset{+1}{N_2O} + 2H_2O$$
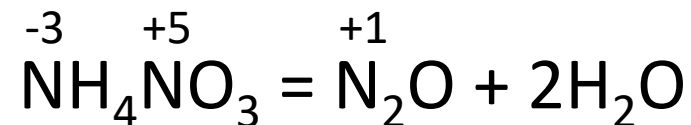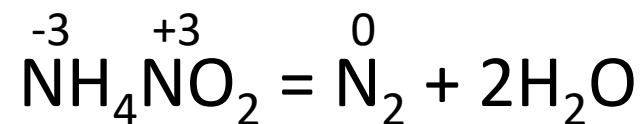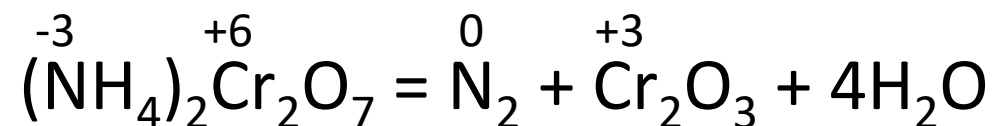
Disproportionation (dismutation) is a type of red-ox reaction in which **one element is simultaneously reduced and oxidized to form two different products.**

$$2H_2\overset{-1}{O}_2 = 2H_2\overset{-2}{O} + \overset{0}{O}_2$$

$$\overset{0}{Br}_2 + 2NaOH = Na\overset{-1}{Br} + Na\overset{+1}{Br}O + H_2O$$

$$4Na\overset{+5}{Cl}O_3 = Na\overset{-1}{Cl} + 3Na\overset{+7}{Cl}O_4$$

$$3K_2\overset{+6}{Mn}O_4 + 2H_2O = 2K\overset{+7}{Mn}O_4 + \overset{+4}{Mn}O_2 + 4KOH$$
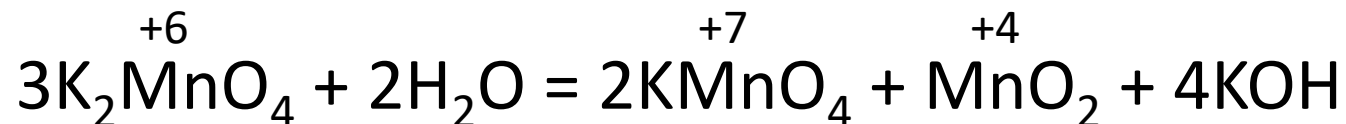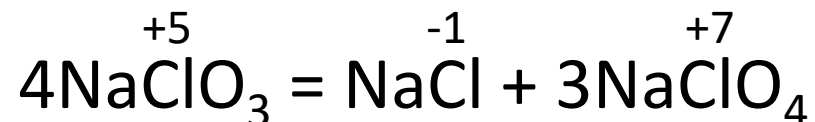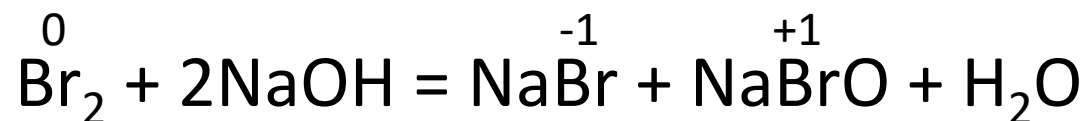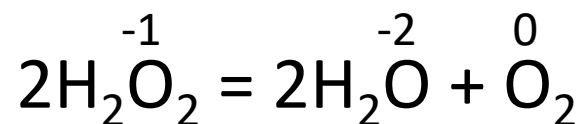
Factors affecting the course of red-ox reactions

Reaction products are change depending on the acidity of the solution, that is, the process of oxidation or reduction will occur in different ways. For example,

The oxidized form

The reduced form

$$+ 8H^+ + 5\bar{e} \longrightarrow Mn^{2+} + 4H_2O$$
colorless
solution

$$MnO_4^- + 2H_2O + 3\bar{e} \longrightarrow MnO_2 + 4OH^-$$
violet
solution

brown
precipitate

$$+\bar{e} \longrightarrow MnO_4^{2-}$$
green
solution
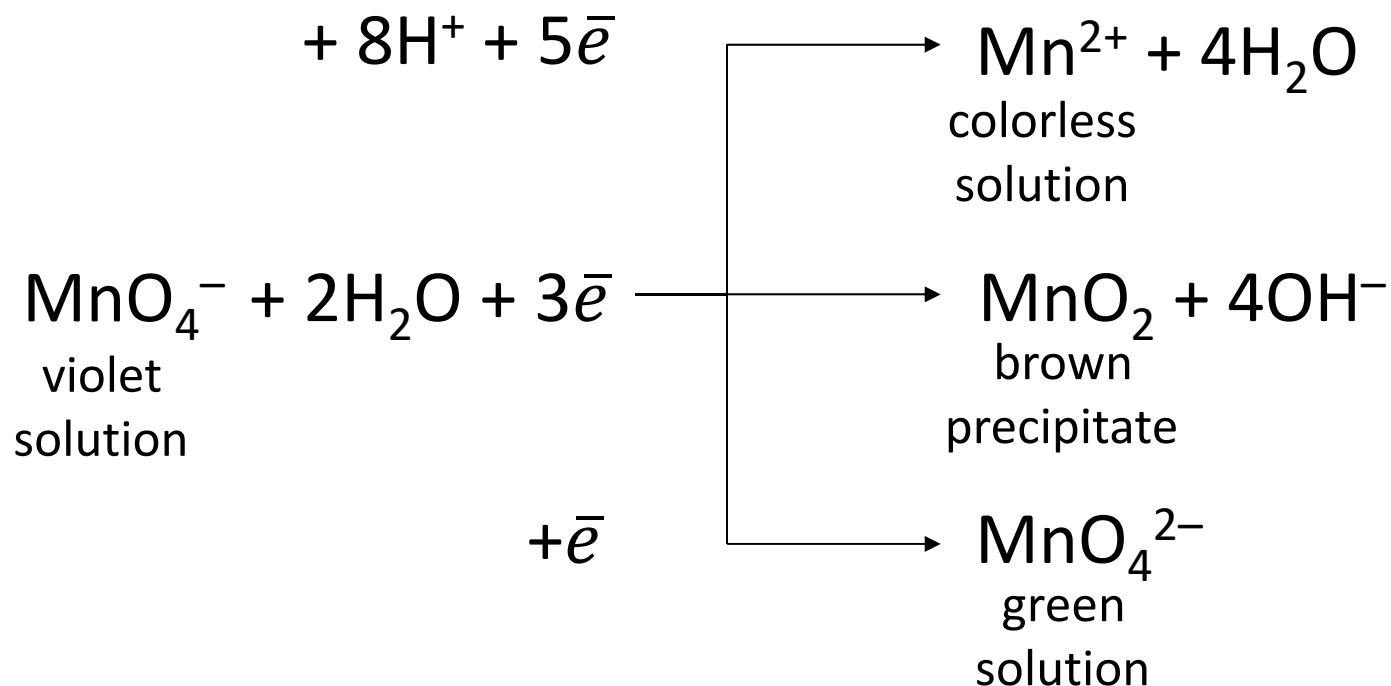
The dilute sulphuric acid is used to create an acidic medium. Nitric acid of any concentration is a strong oxidizing agent, and hydrochloric acid exhibits reducing properties. Alkaline medium is created by adding potassium or sodium hydroxide.

Similarly, the temperature affects on the redox interaction: some reactions occur only with heating or depending on the temperature patterns give different reaction products. For example, at room temperature chlorine disproportionates to form compounds with oxidation states -1 and +1:

$$\overset{0}{Cl_2} + NaOH = Na\overset{-1}{Cl} + Na\overset{+1}{Cl}O + H_2O$$

$$-\bar{e} + Cl^0 + 2OH^- \rightleftarrows ClO^- + H_2O \qquad | \quad 1$$

$$+\bar{e} + Cl^0 + \rightleftarrows Cl^- \qquad | \quad 1$$

$$Cl_2 + 2OH^- = Cl^- + ClO^- + H_2O$$

$$Cl_2 + 2NaOH = NaCl + NaClO + H_2O$$
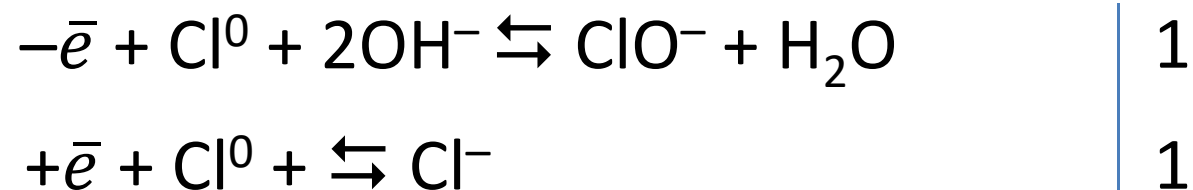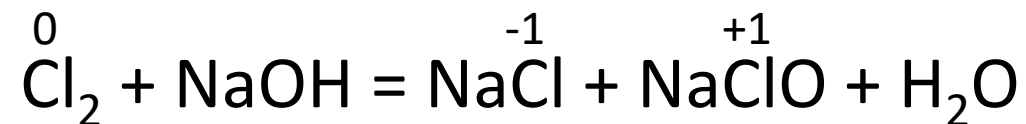
When the same reaction occurs in boiling solution, compounds with oxidation states -1 and +5 are formed:

$$\overset{0}{Cl_2} + NaOH = Na\overset{-1}{Cl} + Na\overset{+5}{Cl}O_3 + H_2O$$

$$-5\bar{e} + Cl^0 + 6OH^- \leftrightarrows ClO_3^- + 3H_2O \quad | \quad 1$$

$$+\bar{e} + Cl^0 + \leftrightarrows Cl^- \quad | \quad 5$$

$$3Cl_2 + 6OH^- = 5Cl^- + ClO_3^- + 3H_2O$$

$$3Cl_2 + 6NaOH = 6NaCl + NaClO_3 + 3H_2O$$
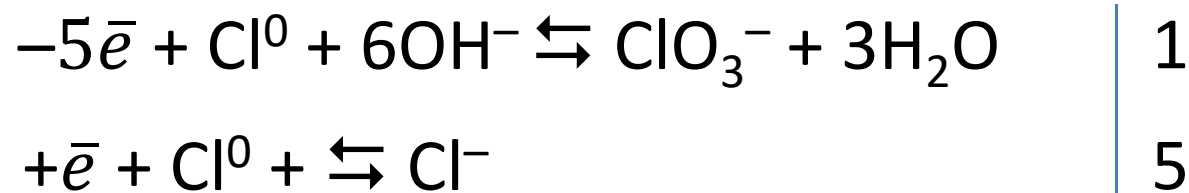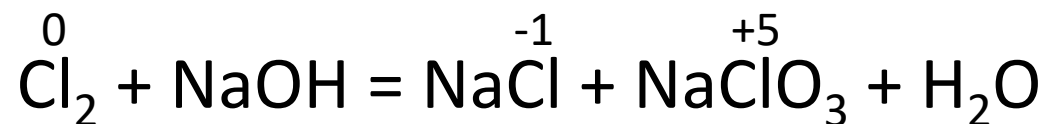
Measurement of Oxidation and Reduction Tendencies

There is a metal strip, immersed in an aqueous solution containing the metal ions, Me^{n+}. A metal strip, Me, called an **electrode**. The entire assembly is called a **half-cell.**
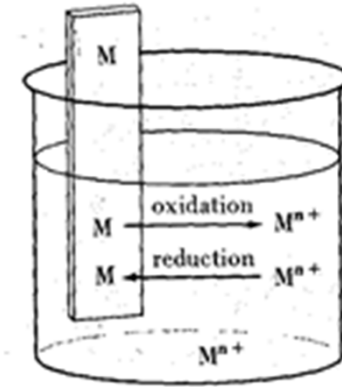
An electrochemical half-cell

Equilibrium between the metal and its ions, which quickly established, can be represented as:

$$Zn_{(s)} \xrightarrow{\text{oxidation}} Zn^{n+}(aq) + n\bar{e}$$
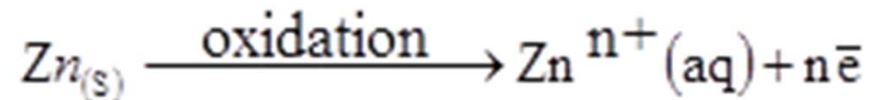
As a result, the electrode acquires a very slight electric change; the solution near the electrode acquires the opposite charge. It is impossible to fix the electrode potential between the metal surface and solution near the electrode.
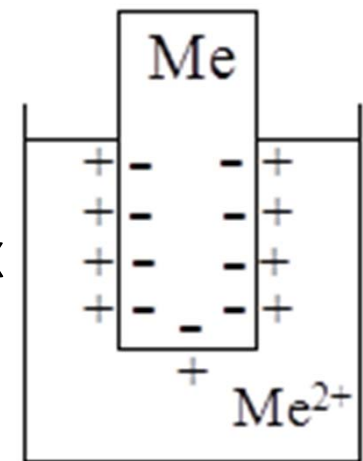
If we have a way to assign numerical values of metal potential, we must compare it to reference electrode (**standard hydrogen electrode**).

The standard hydrogen electrode involves H^+ ions in solution H_2SO_4 with concentration 1M. Hydrogen molecules in the gaseous state are at a pressure of 1atm. The value of the potential standard hydrogen electrode is set at zero.

The standard hydrogen electrode

$$2H^+ + 2\bar{e} \rightleftarrows H_2 \quad E^0 = 0.0000 \text{ Volt (V)}$$
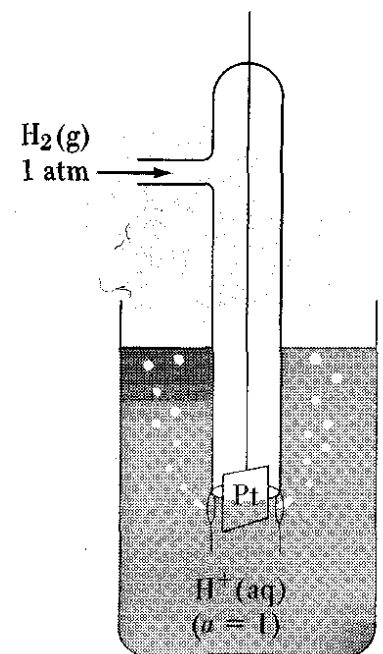
If we connect the metal electrode with the standard hydrogen electrode by the potentiometer and salt bridge, we might be able to calculate cell potentials of this metal.
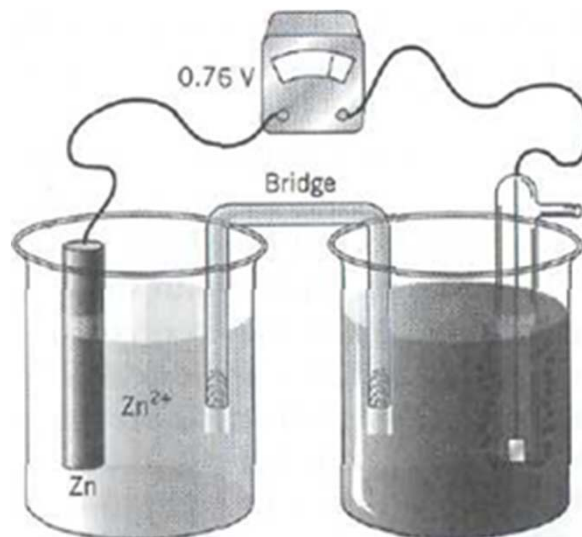
The symbol for standard electrode potential is **E°**. When a standard hydrogen electrode is combined with a standard zinc electrode, electrons are transfered from the zinc to the hydrogen electrode. The standard hydrogen electrode acts as **the cathode** and the standard zinc electrode as **the anode**. A combination of two half-cells is called an **electrochemical cell.**

The reaction that occurs in the cell is:

$$\text{oxidation: } Zn^0 - 2\bar{e} \rightarrow Zn^{2+}$$

$$\text{reduction: } 2H^+ + 2\bar{e} \rightarrow H_2^0$$
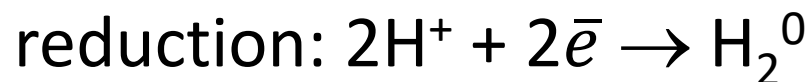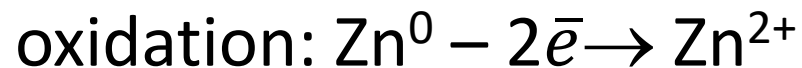
The measured value of standard electrode potential for Zn, $E°$ is $-0.76V$. When a standard hydrogen electrode is combined with a standard copper (Cu^0) electrode, the reaction that occurs in the cell is:

$$\text{oxidation: } H_2^0 - 2\bar{e} \rightarrow 2H^+$$

$$\text{reduction: } Cu^{2+} + 2\bar{e} \rightarrow Cu^0$$
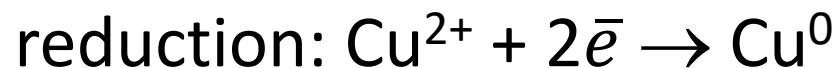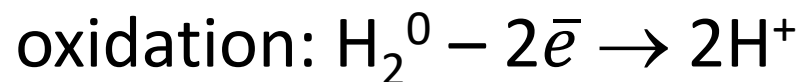
The measured value of standard electrode potential for Cu, $E°$ is $+ 0.34V$

The Activity Series of Metals

   This table shows another redox table that includes some common metals and the hydrogen ion. This table is often called the **Activity Series of Metals** because it lists common metals according to their strength as reducing agents. Stronger reducing agents occur in the left of the list and they decrease in strength from the left to the right of the table.

   Metals are listed in the table as reducing agents; metal ions are listed as oxidizing agents. The table shows which metals and metal ions react with one another. A specific metal ion reacts with any metal that appears below in the table.

Na	Mg	Al	Zn	Sn	H	Cu	Ag	Pt	Au
Na^+/Na	Mg^{2+}/Mg	Al^{3+}/Al	Zn^{2+}/Zn	Sn^{2+}/Sn	$2H^+/H_2$	Cu^{2+}/Cu	Ag^+/Ag	Pt^{2+}/Pt	Au^{3+}/Au
- 2.71	- 2.36	- 1.67	- 0.76	- 0.14	0.00	0.34	0.80	1.19	1.50

Example.

When zinc metal is placed in a solution containing copper (II) ions, Cu^{2+}, a reaction occurs.

$$Zn_{(s)} + Cu^{2+} \rightarrow Zn^{2+} + Cu_{(s)}$$
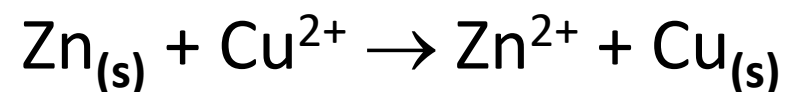
Thus, the predicted reaction is zinc metal reaction with copper (II) ion to give copper metal and zinc ion.

Table also shows why some metals dissolve in acidic solutions and some do not. When any of the metals below H^+ in the table are added to an acidic solution, a reaction occurs producing H_2 gas and the ion of the metal. These metals dissolve in acidic solution. For example, note relative positions of the hydrogen ion and magnesium metal in the table.

$$2H^+ + 2\bar{e} \rightarrow H_{2(g)}$$
$$Mg^{2+} - 2\bar{e} \rightarrow Mg_{(s)}$$
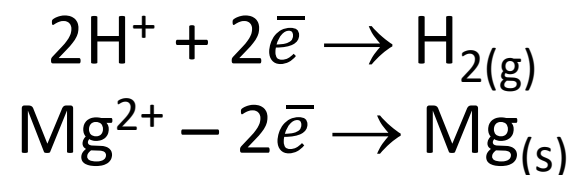
When magnesium metal is added to a solution of hydrochloric acid, a reaction occurs.

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

Note: H_2O also appears in the table as an oxidizing agent. Some more active metals readily react with water. For example, when sodium metal is placed in water, a very exothermic reaction occurs.

$$2Na_{(s)} + 2H_2O \rightarrow 2NaOH_{(aq)} + H_{2(g)}$$
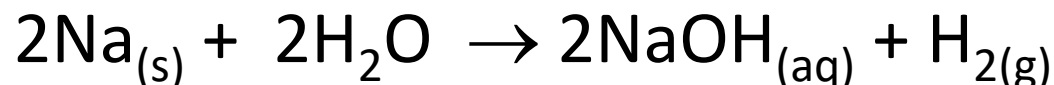
Summary

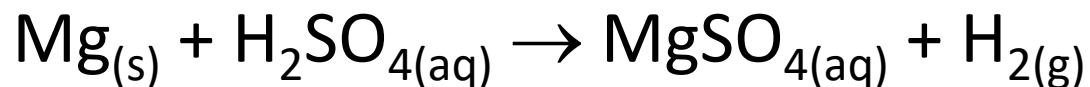If all the metals are arranged in order of increasing of their standard electrode potentials, the so-called **electrochemical potential line of metals**. The following conclusions can be made by the help of it:

The metals, which dispose before the hydrogen in the row of standard electrode potentials dissolve in $HCl_{(aq)}$ and $H_2SO_{4(aq)}$ producing $H_{2(g)}$

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$
$$Mg_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + H_{2(g)}$$

Copper, Mercury and Silver does not dissolve in $HCl_{(aq)}$ and $H_2SO_{4(aq)}$.

The metals which dispose in the left side after Mg, displace the after metals, which dispose in the right of them such as:

$$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$$
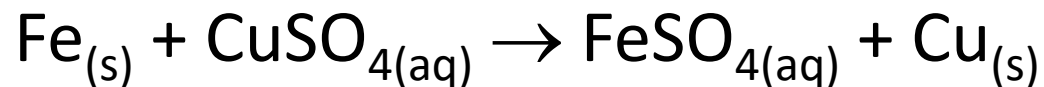
Some standard electrode potentials

Reduction half-reaction	E°, V
$F_{2\,(g)} + 2\bar{e} \rightarrow 2F^-_{(aq)}$	+2.87
$O_{3(g)} + 2H^+_{(aq)} + 2\bar{e} \rightarrow O_{2(g)} + H_2O$	+2.07
$S_2O_8^{2-}_{(aq)} + 2\bar{e} \rightarrow 2SO_4^{2-}$	+2.01
$H_2O_{2(aq)} + 2H^+_{(aq)} + 2\bar{e} \rightarrow 2H_2O$	+1.77
$MnO_4^-_{(aq)} + 8H^+_{(aq)} + 5\bar{e} \rightarrow Mn^{2+}_{(aq)} + 4H_2O$	+1.51
$PbO_{2\,(s)} + 4H^+_{(aq)} + 2\bar{e} \rightarrow Pb^{2+}_{(aq)} + 2H_2O$	+1.455
$Cl_{2(g)} + 2\bar{e} \rightarrow 2Cl^-_{(aq)}$	+1.360
$Cr_2O_7^{2-}_{(aq)} + 14H^+_{(aq)} + 6\bar{e} \rightarrow 2Cr^{3+}_{(aq)} + 7H_2O$	+1.33
$O_{2(g)} + 4H^+_{(aq)} + 4\bar{e} \rightarrow 2H_2O$	+1.32
$Ag^+_{(aq)} + \bar{e} \rightarrow Ag_{(s)}$	+0.80
$Fe^{3+}_{(aq)} + \bar{e} \rightarrow Fe^{2+}_{(aq)}$	+0.771
$O_{2(g)} + 2H^+_{(aq)} + 2\bar{e} \rightarrow H_2O_{2(aq)}$	+0.682
$MnO_4^-_{(aq)} + 2H_2O_{(aq)} + 3\bar{e} \rightarrow MnO_{2\,(s)} + 4OH^-$	+0.57
$MnO_4^-_{(aq)} + \bar{e} \rightarrow MnO_4^{2-}_{(aq)}$	+0.564
$I_{2(s)} + 2\bar{e} \rightarrow 2I^-_{(aq)}$	+0.535
$Cu^{2+}_{(aq)} + 2\bar{e} \rightarrow Cu_{(s)}$	+0.337
$SO_4^{2-}_{(aq)} + 4H^+_{(aq)} + 2\bar{e} \rightarrow 2H_2O + SO_{2(g)}$	+0.17
$\mathbf{2H^+_{(aq)} + 2\bar{e} \rightarrow H_{2(g)}}$	**+0.000**
$Pb^{2+}_{(aq)} + 2\bar{e} \rightarrow Pb_{(s)}$	−0.126
$CrO_4^{2-}_{(aq)} + 4H_2O + 3\bar{e} \rightarrow Cr(OH)_{3(s)} + 5OH^-$	−0.13
$Sn^{2+}_{(aq)} + 2\bar{e} \rightarrow Sn_{(s)}$	−0.136
$Fe^{2+}_{(aq)} + 2\bar{e} \rightarrow Fe_{(s)}$	−0.440
$Zn^{2+}_{(aq)} + 2\bar{e} \rightarrow Zn_{(s)}$	−0.763
$SO_4^{2-}_{(aq)} + H_2O + 2\bar{e} \rightarrow SO_3^{2-}_{(aq)} + 2OH^-$	−0.93
$Al^{3+}_{(aq)} + 3\bar{e} \rightarrow Al_{(s)}$	−1.66
$Mg^{2+}_{(aq)} + 2\bar{e} \rightarrow Mg_{(s)}$	−2.376

Direction of Redox Reactions

The direction of the redox reaction is determined by E^0_{cell} (electromotive force). It is E^0_{cell} of Galvanic cell formed from this oxidation and reduction agents.

$$E^0_{cell} = E_{ox} - E_{red},$$

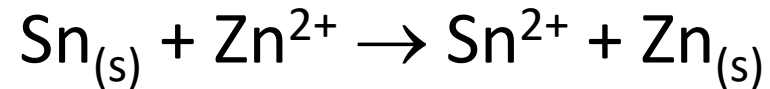where E_{ox}, E_{red} − potentials of the oxidant and reductant, respectively.

If $E^0_{cell} > 0$, that is $E_{ox} > E_{red}$, this reaction is possible, flows in the forward direction.

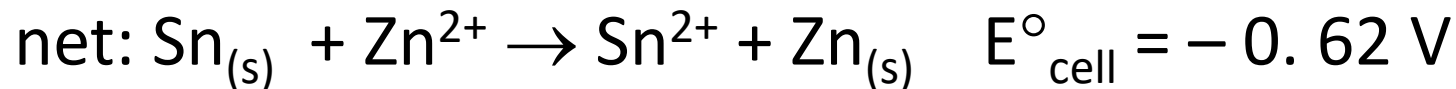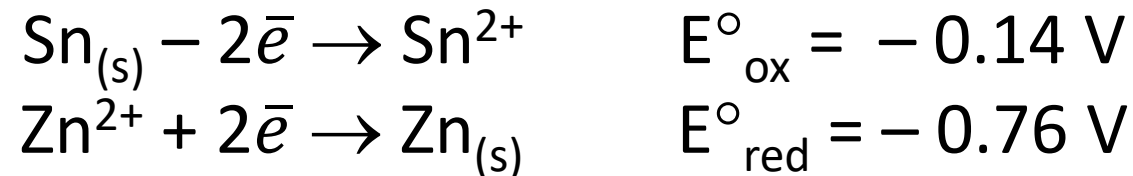If E°_{cell} is *positive,* the reaction will occur spontaneously in the forward direction.

If E°_{cell} is *negative,* the reaction will proceed spontaneously in the reverse direction.

Example

$$Sn_{(s)} + Zn^{2+} \rightarrow Sn^{2+} + Zn_{(s)}$$

Solution

The reaction that occurs in the cell is:

$$Sn_{(s)} - 2\bar{e} \rightarrow Sn^{2+} \qquad E°_{ox} = -0.14 \text{ V}$$
$$Zn^{2+} + 2\bar{e} \rightarrow Zn_{(s)} \qquad E°_{red} = -0.76 \text{ V}$$

net: $Sn_{(s)} + Zn^{2+} \rightarrow Sn^{2+} + Zn_{(s)}$ $E°_{cell} = -0.62 \text{ V}$

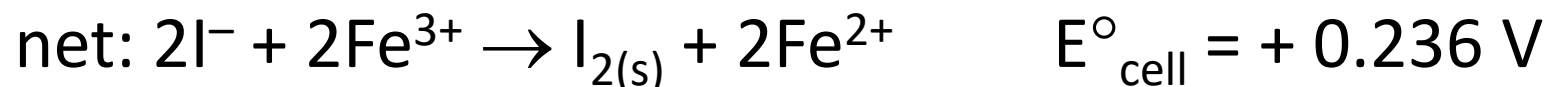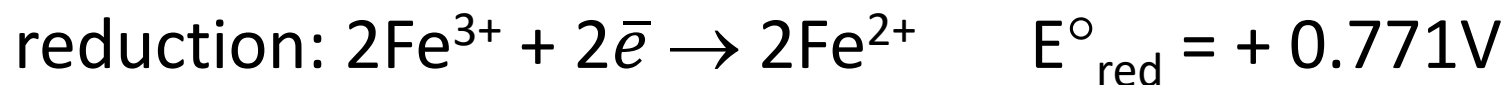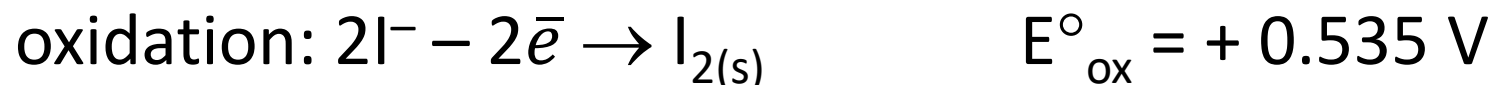**E°$_{cell}$ = E°$_{ox}$ − E°$_{red}$** = –0.76 – (–0.14) = –0.62V

Since E°$_{cell}$ has negative values, the reaction can't go to the products. The $Sn_{(s)}$ can't displace Zn^{2+} from aqueous solution.

Example

$$2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_{2(s)}$$

Solution (see in the some standard electrode potentials). The reaction that occurs in the cell is:

oxidation: $2I^- - 2\bar{e} \rightarrow I_{2(s)}$ $E^\circ_{ox} = +0.535$ V

reduction: $2Fe^{3+} + 2\bar{e} \rightarrow 2Fe^{2+}$ $E^\circ_{red} = +0.771$V

net: $2I^- + 2Fe^{3+} \rightarrow I_{2(s)} + 2Fe^{2+}$ $E^\circ_{cell} = +0.236$ V

$$E^\circ_{cell} = 0.771 - 0.535 = 0.236 \text{ V}$$

Since E°_{cell} has positive values.

We can draw the conduction:

The salts of Fe^{3+} can oxidize the I^- ions to the $I_{2(s)}$.

Example: determine the direction in which the reaction may flows accidentally**:**

$$2NaCl + Fe_2(SO_4)_3 = 2FeSO_4 + Cl_2 + Na_2SO_4$$

Solution

$$2Cl^- - 2\bar{e} \leftrightarrows Cl_2 \qquad\qquad E^0{}_1 = 1.36 \text{ V}$$

$$Fe^{3+} + 2\bar{e} \leftrightarrows Fe^{2+} \qquad\qquad E^0{}_2 = 0.77 \text{ V}$$

$$E^0{}_{cell} = E_{ox} - E_{red} = 0.77 - 1.36 = -0.59V$$

Since $E^\circ{}_{cell}$ has negative values, the reaction can't go to the products.

The salts of Fe^{3+} can't oxidize the Cl^- ions to the $Cl_{2(g)}$. The reaction can take place only from right to the left.

Thank you for attention!