# Oxidation and Reduction Reactions Workbook

## Reference sheets:

The activity series of common metals
Calculating oxidation numbers

## Work sheets

1. Oxidation, Reduction, Agents, & Reactions. WS 1
2. Oxidation Numbers Spontaneous Reactions WS 2
3. Oxidation Numbers, Application to Reactions. WS 3

# The activity series of common metals

*Activity series of elements*

| elements | | | | |
|---|---|---|---|---|
| Lithium | Li | $\rightarrow$ | $Li^+$ | + e- |
| Potassium | K | $\rightarrow$ | $K^+$ | + e- |
| Barium | Ba | $\rightarrow$ | $Ba^{2+}$ | + 2e- |
| Calcium | Ca | $\rightarrow$ | $Ca^{2+}$ | + 2e- |
| Sodium | Na | $\rightarrow$ | $Na^+$ | + e- |
| Magnesium | Mg | $\rightarrow$ | $Mg^{2+}$ | + 2e- |
| Aluminum | Al | $\rightarrow$ | $Al^{3+}$ | + 3e- |
| Manganese | Mn | $\rightarrow$ | $Mn^{2+}$ | + 2e- |
| Zinc | Zn | $\rightarrow$ | $Zn^{2+}$ | + 2e- |
| Chromium | Cr | $\rightarrow$ | $Cr^{3+}$ | + 3e- |
| Iron | Fe | $\rightarrow$ | $Fe^{2+}$ | + 2e- |
| Cobalt | Co | $\rightarrow$ | $Co^{2+}$ | + 2e- |
| Nickel | Ni | $\rightarrow$ | $Ni^{2+}$ | + 2e- |
| Tin | Sn | $\rightarrow$ | $Sn^{2+}$ | + 2e- |
| Lead | Pb | $\rightarrow$ | $Pb^{2+}$ | + 2e- |
| HYDROGEN | $H_2$ | $\rightarrow$ | $2H^+$ | + 2e- |
| Copper | Cu | $\rightarrow$ | $Cu^{2+}$ | + 2e- |
| Mercury | 2Hg | $\rightarrow$ | $Hg_2^{2+}$ | + 2e- |
| Silver | Ag | $\rightarrow$ | $Ag^+$ | + e- |
| Mercury | Hg | $\rightarrow$ | $Hg^{2+}$ | + 2e- |
| Platinum | Pt | $\rightarrow$ | $Pt^{2+}$ | + 2e- |
| Gold | Au | $\rightarrow$ | $Au^{3+}$ | + 3e- |

- Elements that lie near the top of the list are referred to as active metals.
- Elements that lie near the bottom of the activity series are very stable and form compounds less readily. Notice, also, that the transition elements from groups 8B to 1B are near the bottom of the list. The alkali and alkaline earth metals are at the top. They are most easily oxidized.
- Any metal on the list can be oxidized by the ions of elements below it. Example: Copper is above silver in the series. Cu metal can be oxidized by Ag+ to give silver metal and copper ions.
- The following react vigorously with acidic solutions to give hydrogen gas and cations of the metals, and hydroxide: Li, K, Ba, Ca, Na.
- The following react vigorously with water to give hydrogen gas and cations of the metals, and hydroxide: Li, K, Ba, Ca, Na.
- The following react with acid to give hydrogen gas and cations of the metal, but not vigorously: Mg, Al, Zn, Cr, Fe, Cd, Co, Ni, Sn, Pb.
- The following react slowly with water but readily with steam to give hydrogen gas and cations: Mg, Al, Zn, Cr, Fe, Cd.
- The following do not react with acids (HCl, HBr, HI) to give hydrogen: Cu, Hg, Ag, Au, Pt.

# Calculating oxidation numbers

Oxidation numbers are bookkeeping numbers. They mark the flow of electrons and are useful for balancing redox (*red*uction/*ox*idation) equations. Oxidation numbers are positive or negative numbers, but are not the same as valance values or the actual charge on ions or atoms.

## General rules: they always apply

Oxidation numbers are assigned to elements using these rules:

- **Rule 1:** The oxidation number of an element in its free (uncombined) state is zero — for example, Al(s) or Zn(s). This is also true for elements found in nature as *diatomic* (two-atom) elements: $H_2$, $O_2$, $S_8$.
- **Rule 2:** The oxidation number of a *monatomic* (one-atom) ion is the same as the charge on the ion, for example: $Na^+$, $S^{2-}$
- **Rule 3:** The sum of all oxidation numbers in a neutral compound is zero. The sum of all oxidation numbers in a *polyatomic* (many-atom) ion is equal to the charge on the ion. This rule often allows chemists to calculate the oxidation number of an atom that may have multiple oxidation states, if the other atoms in the ion have known oxidation numbers.
- **Rule 4:** The oxidation number of an alkali metal (IA family) in a compound is +1; the oxidation number of an alkaline earth metal (IIA family) in a compound is +2.
- **Rule 5:** The oxidation number of oxygen in a compound is usually –2. If, however, the oxygen is in a class of compounds called *peroxides* (for example, hydrogen peroxide), then the oxygen has an oxidation number of –1. If the oxygen is bonded to fluorine, the number is +1.
- **Rule 6:** The oxidation state of hydrogen in a compound is usually +1. If the hydrogen is part of a *binary metal hydride* (compound of hydrogen and some metal), then the oxidation state of hydrogen is –1.
- **Rule 7:** The oxidation number of fluorine is always –1. Chlorine, bromine, and iodine usually have an oxidation number of –1, unless they're in combination with an oxygen or fluorine.

These rules give you another way to define oxidation and reduction — in terms of oxidation numbers. For example, consider this reaction, which shows oxidation by the loss of electrons:

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\ e^-$$

Notice that the zinc metal (the reactant) has an oxidation number of zero **(rule 1**), and the zinc cation (the product) has an oxidation number of +2 (rule 2). In general, you can say that a substance is oxidized when there's an increase in its oxidation number.

Reduction works the same way. Consider this reaction:

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

The copper is going from an oxidation number of +2 to zero. A substance is reduced if there's a decrease in its oxidation number. The rules below are used to predict oxidation numbers.

## Special Rules: read carefully!

These rules only apply to certain columns on the Periodic Table. They often combine with the general rules. You might need a periodic table as a reference.

The first rules apply to the main group elements. Different Periodic Tables label these in different ways, so two types of numbers are given. The currently accepted method is to use the numbers 1-2 and 13-18 for the representative elements. Older tables used IA-VIIIA (or IB -VIIIB).

Each of the rules will use both methods with the newer method of labeling given in parentheses. Note: These are the most common/stable oxidation numbers. There are exceptions in addition to the ones given below. These values refer to these elements in compounds. General Rule 1 still applies to the elements in the elemental state.

## Representative elements (The A's)

1. Column IA (1) +1 (Except H with a metal, then H = -1)
2. Column IIA (2) +2
3. Column IIIA (13) +3 (Normally, however +1 possible near the bottom of the table)
4. Column IVA (14) +4 to -4
5. Column VA (15) +5 to -3
6. Column VIA (16) +6 to -2 (Oxygen is -2 except when combined with F, or in $O_2$ or $O_2^{2-}$)
7. Column VIIA (17) +7 to -1 (Fluorine is -1 in compounds)
8. Column VIIIA (18) +8 to 0 (Usually only 0)

## Transition elements (The B's)

Transition Metals IIIB-IIB (3-12). If the B designation is used the values may range from +2 to the group number. If the other designation is used, the values may range from +2 to the group number for columns 3-8 and to the group number minus 10 for groups 11 and 12 (treat columns 9 and 10 as if they were 7 and 6 respectively). Exceptions: $Hg_2^{2+}$ (Hg = +1) and $Au^{3+}$ (Au = +3).

When an element may have a positive or negative oxidation number, it will normally be negative if it is to the upper right, on the Periodic Table, with respect to the other elements in the compound or ion.

When it is negative, under these circumstances, it will probably have the most negative of the possible values noted in the rules above. Note that in Special Rules 4-8 the range of possible values is always eight. With these elements, the
more probable oxidation states may be determined by counting from highest to lowest by twos.

**Worksheet #1** **Writing half reactions**

1. Define each: Remember "Oil Rig": Oxidation is loss (of $e^-$)reduction is gain (of $e^-$)

   a) Oxidation

   b) Reduction

   c) Oxidizing agent

   d) Reducing agent

2. Write half reactions for each of the following atoms or ions. Label each as oxidation or reduction. Use the Activity Series Chart.

   a) Al

   b) $Ba^{2+}$

   c) $Br_2$

   d) Ca

   e) $Ga^{3+}$

   f) $H_2$

   g) $H^+$

3. Balance the spontaneous redox reaction below. A spontaneous reaction is a reaction that occurs: 1) by a driving force that favors the product, 2) the free energy of the product is lower than the free energy of the reactant, and/or 3) occurs without any outside 'help' such as electrolysis. Identify the entities reduced and oxidized. State the reducing agent and the oxidizing agent.

   a) $Al + Zn^{2+}$

   b) $F_2 + O^{2-}$

   c) $O_2 + Ca$

   d) $Al^{3+} + Li$

4. Write the oxidation and reduction reactions for each redox reaction. The first one is done for you.

   a) $Fe^{2+}_{(aq)} + Co_{(s)} \rightarrow Co^{2+}_{(aq)} + Fe_{(s)}$

      (i) Oxidation: $Co_{(s)} \rightarrow Co^{2+}_{(aq)} + 2e^-$

      (ii) Reduction: $Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$

   b) $Ag^{+}_{(aq)} + Ni_{(s)} \rightarrow Ni^{3+}_{(aq)} + 3\ Ag_{(s)}$

      (i) Oxidation:

      (ii) Reduction:

   c) $Cu^{2+}_{(aq)} + Pb_{(s)} \rightarrow Cu^{2+}_{(aq)} + Pb_{(s)}$

      (i) Oxidation:

      (ii) Reduction:

   d) $Sn_{(s)} + O_{2(g)} \rightarrow O^{2-} + 2Sn^{2+}_{(aq)}$

      (i) Oxidation:

      (ii) Reduction:

   e) $Co^{2+}_{(aq)} + 2\ F^{-}_{(aq)} \rightarrow Co_{(s)} + F_{2(g)}$

      i) Oxidation:

      ii) Reduction:

**Worksheet #2** **Finding oxidation numbers**

1. Determine the Oxidation Number of each of the elements that is underlined.

a) $\underline{N}H_3$ _____ b) $H_2\underline{S}O_4$ _____

c) $Zn\underline{S}O_3$ _____ d) $\underline{Al}(OH)_3$ _____

e) $\underline{Na}$ _____ f) $\underline{Cl}_2$ _____

g) $Ag\underline{N}O_3$ _____ h) $\underline{Cl}O_4^-$ _____

i) $\underline{S}O_2$ _____ j) $K_2\underline{Cr}_2O_4$ _____

k) $Ca(\underline{Cl}O_3)_2$ _____ l) $K_2\underline{Cr}_2O_7$ _____

m) $H\underline{P}O_3^{2-}$ _____ n) $H\underline{Cl}O$ _____

o) $\underline{Mn}O_2$ _____ p) $K\underline{Cl}O_3$ _____

q) $\underline{Pb}O_2$ _____ r) $\underline{Pb}SO_4$ _____

s) $K_2\underline{S}O_4$ _____ t) $\underline{N}H_4^+$ _____

u) $Na_2\underline{O}_2$ _____ v) $\underline{Fe}O$ _____

w) $\underline{Fe}_2O_3$ _____ x) $Si\underline{O}_4^{4-}$ _____

y) $Na\underline{I}O_3$ _____ z) $\underline{Cl}O_3^-$ _____

aa) $\underline{N}O_3^-$ _____ bb) $\underline{Cr}(OH)_4$ _____

cc) $Ca\underline{H}_2$ _____ dd) $\underline{Pt}(H_2O)_5(OH)^{2+}$ _____

ee) $Fe(H_2O)_6^{3+}$ _____ ff) $CH_3COOH$ _____

2. What is the oxidation number of carbon in each of the following substances?

a) CO _____ b) C _____

c) $CO_2$ _____ d) $CO_3^{2-}$ _____

e) $C_2H_6$ _____ f) $CH_3OH$ _____

3. For each of the following reactants, identify: the oxidizing agent, the reducing agent, the substance oxidized and the substance reduced.

a) $Cu^{2+}{}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}{}_{(aq)}$

Substance oxidized _____ Substance reduced _____
Oxidizing agent _____ Reducing agent _____

b) $Cl_{2\,(g)} + 2\,Na_{(s)} \rightarrow 2\,Na^{+}{}_{(aq)} + 2\,Cl^{-}{}_{(aq)}$

Substance oxidized _____ Substance reduced ____
Oxidizing agent _____ Reducing agent _____

**Worksheet # 3**
**Decide if the reaction will go (Spontaneous and Non-spontaneous) Redox Reactions**

1. Describe each reaction as spontaneous or non-spontaneous.

   a) $Au^{+3}+Fe^{+3}\rightarrow Fe^{+2}+Au$

   b) $Pb+Fe\rightarrow Fe^{+2}+Pb^{+2}$

   c) $Cl_2+2F^{-}\rightarrow F_2+2Cl^{-}$

   d) $Mg^{-2}+Pb\rightarrow Mg+Pb^{+2}$

   e) $2I^{-}+Cl_2\rightarrow 2Cl^{-}+I_2$

   **f)** $Pb^{+2}+Au\rightarrow Au^{+3}+Pb$

2. Can you keep 1 M HCl in an iron container? If the answer is no, write a balanced equation for the reaction that would occur.

3. Can you keep 1 M HCl in an Ag container? If the answer is no, write a balanced equation for the reaction that would occur.

4. Can you keep 1 M $HNO_3$ in an Au container? If the answer is no, write a balanced equation for the reaction that would occur. (Remember, $HNO_3$ consists of two ions $H^+$ and $NO_3^-$)

5. Circle each formula that can lose an electron

   $O_2$ $Cl^{-}$ Fe $Na^{+}$

4. Determine the oxidation number for the element underlined

| | | | |
|---|---|---|---|
| $Pb\underline{S}O_4$ | _____ | $\underline{Cl}O_3^-$ | _____ |
| $H\underline{P}0_3^{2-}$ | _____ | $Na_2\underline{O}_2$ | _____ |
| $Ca\underline{H}_2$ | _____ | $Al_2(\underline{S}O_4)_3$ | _____ |
| $Na\underline{I}O_3$ | _____ | $\underline{C}_4H_{12}$ | _____ |

5. For each of the following reactants, identify: the oxidizing agent, the reducing agent, the substance oxidized and the substance reduced. The equations are NOT balanced-don't balance them

   $Al^{3+}+Zn\rightarrow Al+Zn^{2+}$

Substance oxidized _____ Oxidizing agent _____

$Cr_2O_7^{2-} + ClO_2^- \rightarrow Cr^{3+} + ClO_4^-$

Substance reduced _____ Oxidizing agent _____

$O_3 + H_2O + SO_2 \rightarrow SO_4^{2-} + O_2 + 2H^+$

Substance oxidized _____ Reducing agent _____

$As_2O_3 + NO_3^- + H_2O + H^+ \rightarrow H_3AsO_4 + NO$

Substance reduced _____ Reducing agent_____

6. Circle each formula that is able to lose an electron

$O_2$ $Cl^-$ Fe $Na^+$

## Redox Half Reactions and Reactions WS #1

Define each

- (g) Oxidation **- loss of electrons**
- (g) Reduction **- gain of electrons**
- (g) Oxidizing agent **- causes oxidation by undergoing reduction**
- (g) Reducing agent **- causes reduction by undergoing oxidation**

Write half reactions for each of the following atoms or ions. Label each as oxidation or reduction.

1. **$Al \rightarrow Al^{3} + 3e^{-}$ oxidation**
2. **$Ba^{2+} + 2e^{-} \rightarrow Ba$ reduction**
3. **$Ca \rightarrow Ca^{2+} + 2e^{-}$ oxidation**
4. **$Ga^{3+} + 3e \rightarrow Ga$ reduction**
5. **$H_2 \rightarrow 2H^{+} + 2e^{-}$ oxidation**
6. **$2H^{+} + 2e \rightarrow H_2$ reduction**

Balance each spontaneous redox equation. Identify the entities reduced and oxidized. State the reducing agent and the oxidizing agent.

1. Al & $Zn^{2+}$

   i. **$2Al + 3Zn^{2+} \rightarrow 2Al^{3+} + 3Zn$**
   ii. **oxidized reduced**
   iii. **reducing agent oxidizing agent**

2. $F_2$ & $O^{2-}$

   i. **$2F_2 + 2O^{2-} \rightarrow 4F^{-} + O_2$**
   ii. **reduced oxidized**
   iii. **oxidizing agent reducing agent**

3. $O_2$ & Ca

   i. **$2Ca + O_2 \rightarrow 2Ca^{2+} + 2O^{2-}$**
   ii. **oxidized reduced**
   iii. **reducing agent oxidizing agent**

4. $Al^{3+}$ & Li

**$Al^{3+}$ + 3Li → Al + $3Li^{+}$**

**reduced oxidized**

**oxidizing agent reducing agent**

Label the species that is **reduced**, that is **oxidized**, the **reducing agent** and the **oxidizing agent**.

1. $Fe^{2+}$ + Co → $Co^{2+}$ + Fe

   **Co → $Co^{2+}$ + 2e- oxidation $Fe^{2+}$ + 2e- → Fe reduction**

2. 3 $Ag^{+}$ + Ni → $Ni^{3+}$ + 3 Ag

   **Ni → $Ni^{2+}$ + 2e- oxidation $Ag^{+}$ + 1e- → Ag reduction**

3. $Cu^{2+}$ + Pb → $Pb^{2+}$ + Cu

   **Pb → $Pb^{2+}$ + 2e- oxidation $Cu^{2+}$ + 2e- → Cu reduction**

4. $O_2$ + 2 Sn → $O^{2-}$ + 2 $Sn^{2+}$

   **Sn → $Sn^{2+}$ + 2e- oxidation $O_2$ + 4e- → $2O^{2-}$**

   **reduction**

5. $Co^{2+}$ + 2 $F^{-}$ → Co + $F_2$

   **$2F^{-}$ → $F_2$ + 2e- oxidation $Co^{2+}$ + 2e- → Co reduction**

**Redox Half Reactions and Reactions WS #2**

(g) State the Oxidation Number of each of the elements that is underlined.

| | | | |
|---|---|---|---|
| a) $\underline{N}H_3$ | **-3** | b) $H_2\underline{S}O_4$ | **6** |
| c) $Zn\underline{S}O_3$ | **4** | d) $\underline{Al}(OH)_3$ | **3** |
| e) $\underline{Na}$ | **0** | f) $\underline{Cl_2}$ | **0** |
| g) $Ag\underline{N}O_3$ | **5** | h) $\underline{Cl}O_4^{-}$ | **7** |
| i) $\underline{S}O_2$ | **4** | j) $K_2\underline{Cr_2}O_4$ | **3** |
| k) $Ca(\underline{Cl}O_3)_2$ | **5** | l) $K_2\underline{Cr_2}O_7$ | **6** |
| m) $H\underline{P}O_3^{2-}$ | **3** | n) $H\underline{Cl}O$ | **1** |
| o) $\underline{Mn}O_2$ | **4** | p) $K\underline{Cl}O_3$ | **5** |
| q) $\underline{Pb}O_2$ | **4** | r) $\underline{Pb}SO_4$ | **2** |
| s) $K_2\underline{S}O_4$ | **6** | t) $\underline{N}H_4^{+}$ | **-3** |
| u) $Na_2\underline{O_2}$ | **-1** | v) $\underline{Fe}O$ | **2** |
| w) $\underline{Fe_2}O_3$ | **3** | x) $Si\underline{O_4}^{4-}$ | **-2** |
| y) $Na\underline{I}O_3$ | **5** | z) $\underline{Cl}O_3^{-}$ | **5** |

aa) $\underline{N}O_3^-$ **5** bb) $\underline{Cr}(OH)_4$ **4**

cc) $\underline{Ca}H_2$ **-1** dd) $\underline{Pt}(H_20)_5(0H)^{2+}$ **+3**

ee) $\underline{Fe}(H_2O)_6^{3+}$ **+3** ff) $\underline{C}H_3COOH$ **0**

(g) What is the oxidation number of carbon in each of the following substances?

a) CO **2** b) C **0**

c) $CO_2$ **4** d) $CO_3^{2-}$ **4**

e) $C_2H_6$ **-3** f) $CH_3OH$ **-2**

(g) For each of the following reactions, identify: the oxidizing agent, the reducing agent, the substance oxidized and the substance reduced.

a) $Cu^{2+}{}_{(aq)}$ + $Zn_{(s)}$ --------> $Cu_{(s)}$ + $Zn^{2+}{}_{(aq)}$

Substance oxidized **Zn** Substance reduced **$Cu^{2+}$**

Oxidizing agent **$Cu^{2+}$** Reducing agent **Zn**

b) $Cl_{2\,(g)}$ + $2\,Na_{(s)}$ --------> $2\,Na+_{(aq)}$ + $2\,Cl^-{}_{(aq)}$

Substance oxidized **Na** Substance reduced **$Cl_2$**

Oxidizing agent **$Cl_2$** Reducing agent **Na**

## WS # 3 Spontaneous and Non-spontaneous Redox Reactions

Describe each reaction as spontaneous or non-spontaneous.

1. $Au^{+3}$ + $Fe^{+3}$ -----> $Fe^{+2}$ + Au **nonspontaneous (two oxidizing agents)**

2. Pb + $Fe^{+3}$ ------> $Fe^{+2}$ + $Pb^{+2}$ **spontaneous**

3. $Cl_2$ + $F^-$ ------> $F_2$ + $2Cl^-$ **nonspontaneous (see exp 9)**

$2I^- + Cl_2 \rightarrow 2Cl^- + I_2$ **spontaneous(see exp 9)**

4. $Pb^{+2}$ + $Fe^{+2}$ ------> $Fe^{+3}$ + Pb **nonspontaneous**

Can you keep 1 M HCl in an iron container. If the answer is no, write a balanced equation for the reaction that would occur. **No, iron is above hydrogen in the activity series. The metal will react with acid**

$$Fe + 2H^+ \longrightarrow Fe^{2+} + H_2$$

Can you keep 1 M HCl in an Ag container. If the answer is no, write a balanced equation for the reaction that would occur.

**Yes. There is no reaction. Silver is below hydrogen in the activity series**

Can you keep 1 M $HNO_3$ in an Au container. If the answer is no, write a balanced equation for the reaction that would occur. (Remember, $HNO_3$ consists of two ions $H^+$ and $NO_3^-$)

**Yes. There is no reaction.**

Circle each formula that is able to lose an elecron

$O_2$ **$Cl^-$** **Fe** $Na^+$

Determine the oxidation number for the element underlined.

| | | | |
|---|---|---|---|
| $Pb\underline{S}O_4$ | **6** | $\underline{Cl}O_3^-$ | **5** |
| $H\underline{P}0_3^{2-}$ | **3** | $Na_2\underline{O}_2$ | **-1** |
| $Ca\underline{H}_2$ | **-1** | $Al_2(\underline{S}O_4)_3$ | **6** |
| $Na\underline{I}O_3$ | **5** | $\underline{C}_4H_{12}$ | **-3** |

$$Al^{3+} + Zn \longrightarrow Al + Zn^{2+}$$

Substance oxidized **Zn** Oxidizing agent **$Al^{3+}$**

$$Cr_2O_7^{2-} + ClO_2^- \longrightarrow Cr^{3+} + ClO_4^-$$

Substance reduced **$Cr_2O_7^{2-}$** Oxidizing agent **$Cr_2O_7^{2-}$**

State the Oxidation Number of each of the elements that is underlined.

| | | | |
|---|---|---|---|
| a) $\underline{N}H_3$ | **-3** | b) $H_2SO_4$ | **6** |
| c) $Zn\underline{C}O_3$ | **4** | d) $Al(OH)_3$ | **3** |
| e) $\underline{Na}$ | **0** | f) $Cl_2$ | **0** |

$1_{(g)}$ Balance the redox equation using the half reaction method.

**$Al$ + $3Ag^{+}$ ----------> $Al^{3+}$ + $3Ag$**

$1_{(g)}$ Circle each formula that is able to lose an electron

$O_2$ **$Cl^{-}$** **$Fe$** $Na^{+}$

$Al^{3+}$ + $Zn$ ----------> $Al$ + $Zn^{2+}$

Substance oxidized **$Zn$** Oxidizing agent **$Al^{3+}$**

$Cr_2O_7^{2-}$ + $ClO_2^{-}$ ----------> $Cr^{3+}$ + $ClO_4^{-}$

Substance reduced **$Cr_2O_7^{2-}$** Oxidizing agent **$Cr_2O_7^{2-}$**

$O_3$ + $H_2O$ + $SO_2$ -----> $SO_4^{2-}$ + $O_2$ + $2H^{+}$

Substance oxidized **$SO_2$** Reducing agent **$SO_2$**

$3As_2O_3$ + $4NO_3^{-}$ + $7H_2O$ + $4\ H^{+}$ --------> $6H_3AsO_4$ + $4NO$

Substance reduced **$NO_3^{-}$** Reducing agent **$As_2O_3$**