REDOX REACTION

(Reduction-Oxidation Reaction)

What will I learn?

- What is a redox reaction?
- What is oxidation and reduction? (and how to identify them)
- What are oxidizing and reducing agents?
- What is oxidation state?
- What is oxidation and reduction in terms of oxidation state?
- How to assign the oxidation state?
- How to use oxidation state to determine which species is oxidized / reduced?

Redox

Oxidation	Reduction
Gain in oxygen	Loss of oxygen
Loss of hydrogen	Gain in hydrogen
Loss of electrons	Gain of electrons

Redox Reactions

The term **redox** is a shortened form of *oxidation-reduction.*

Oxidation is a process in which one atom loses or donates electrons to another.

Reduction is a process in which one atom gains or accepts electrons from another.

In any redox reaction, one species is ALWAYS oxidized and one is ALWAYS reduced. You cannot have oxidation without reduction also occurring.

Development of oxidation and reduction reaction concept

1. Reaction of reduction oxidation based on releasing (lossing) and gaining of oxygen

a. Oxidation reaction

Oxidation reaction is a reaction of gaining (capturing) of oxygen by a substance

Example :

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{g)}$$

$$P_{4(s)} + 5O_{2(g)} \longrightarrow 2P_2O_{5(s)}$$

b. **Reduction reaction**

Reduction reaction is a reaction of releasing (lossing) of oxygen from an oxide compound

Example:

$$CuO_{(s)} + H_{2(g)} \longrightarrow Cu_{(s)} + H_2O_{(g)}$$

$$Fe_2O_{3(s)} + 3CO_{(g)} \longrightarrow 2Fe_{(s)} + 3CO_{2(g)}$$

2. Reduction oxidation reaction based on electron transfer

a. *Oxidation reaction*

Oxidation reaction is a reaction of **electron releasing** (**lossing)** from a substance.

Example:

$$Na \longrightarrow Na^{+} + e^{-}$$

$$Mg \longrightarrow Mg^{2+} + 2\,e^{-}$$

$$Cu \longrightarrow Cu^{2+} + 2\,e^{-}$$

b. **Reduction reaction**

Reduction reaction is a reaction of **electron gaining** by a substance.

Example:

$$Cl_2 + 2e^{-} \longrightarrow 2Cl^{-}$$

$$S + 2\,e^{-} \longrightarrow S^{2-}$$

3. Reduction oxidation reaction based on oxidation number change

a. *Oxidation reaction*

Oxidation reaction is a chemical reaction which is accompanied by **increasing of oxidation number.**

Example:

$$Al_{(s)} \longrightarrow Al^{3+}_{(aq)}$$

$$S^{2-}_{(aq)} \longrightarrow S_{(s)}$$

b. Reduction reaction

Reduction reaction is a chemical reaction which is accompanied by **decreasing of oxidation number.**

Example:

$$Sn^{4+}_{(aq)} \longrightarrow Sn^{2+}_{(aq)}$$

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

Rules for Assigning Oxidation Numbers

1. The oxidation state of any neutral element in its *naturally occurring state* is zero.

   Na, Be, K, Cl_2, H_2, O_2, P_4 = 0

2. The oxidation number of any cation or anion composed of just one atom is that ion's actual charge.

   H^+= +1; Li^+ = +1; Fe^{3+}= +3; O^{2-} = -2

3. The oxidation number of oxygen is **usually** –2. In H_2O_2 and O_2^{2-} it is –1.

4. The oxidation number of hydrogen is +1 *except* when it is bonded to metals in binary compounds. In these cases, its oxidation number is –1.

Rules for Assigning Oxidation Numbers *continued*

5. The oxidation number of F is *always* -1.

6. Oxidation numbers do not have to be integers. Oxidation number of oxygen in the superoxide ion, O_2^-, is -½.

7. When you begin, assign the most electronegative element present the charge it would have if it were an anion. If oxygen is present, it has the higher priority and will always be –2.

8. ***The sum of the oxidation numbers of all the atoms in a molecule or ion is equal to the charge on the molecule or ion.***

Oxidation Number

is **a number** that states **electrical charge possessed by each one element atom** in **the molecular compound or the ion.**

In the **molecules of ionic compound**, electrical charge contained element atom **can be raised** by **transfering of electrons**.
In the formation of ionic bond:
-Metal atom losses electron to form the positive ion.
-Nonmetal atom gains electron to form the negative ion.
In the molecule of **MgF_2**, consist of **Mg^{2+} ion** with charge of **2+** and **F^- ion** with charge of **1−**
Said that in the molecule of **MgF_2**, **oxidation number of Mg** is **+2**, and **oxidation number** of **F** is **-1**.
In the molecule of covalent compound, the raising of the electrical charge each element atom is caused by its existence the difference of electronegativity of element, so that occur polarization covalent bond.
In the polar covalent compound, the more electronegative atom become more negative charge and the other atom become more positive charge.
In the polar covalent compound of **H_2O, H** contain **1+** and **O** contain **2−**

Oxidation Number Rules

1. ***Oxidation number of free elements***
   Free elements (include molecular elements: H_2, O_2, O_3, N_2, F_2, Cl_2, Br_2, I_2, P_4, S_8) have oxidation number of 0 (zero).
2. ***Oxidation number of fluorine***
   In its compounds, oxidation number of **F** always –**1**.

3. ***Oxidation number of hydrogen***
   In its compounds, oxidation number of **H** always **+1.**
   Except, hydrogen in the **hydride** compounds (compound of H with metal), oxidation number of **H,** is –**1**
   Example:
   In the compound of H_2O, NH_3, H_2S, HCl, HNO_3, H_2SO_4, oxidation number of **H,** is **+1**
   In the **hydride compound,** like LiH, NaH, MgH_2, oxidation number of **H,** is –**1**

4. *Oxidation number of oxygen*

In its compound,oxidation number of **O** always **-2**

For example: compound of Na_2O, CaO, Al_2O_3 , SO_3, CO_2, P_2O_5, Cl_2O_7, H_2CO_3, H_3PO_4, etc, oxidation number of **O,** is -2

Exception occur in the compound of **peroxide, superoxide,** and **F_2O.**

a. In compound of **peroxide,** like Na_2O_2, BaO_2, H_2O_2, oxidation number of O, is –**1.**

b. In compound of **superoxide,** KO_2, oxidation number of O, is –½**.**

c. In compound of **F_2O**, oxidation number of O, is **+2**.

5. *Oxidation number of metals*

Metallic elements in its compound always has oxidation number with **positive** sign.

a. Metallic elements of group A in its compound only has one type of oxidation number, that is the same with its group number, **except** metallic elements of **group IVA and VA**

Metals of group **IA** (Alkali metals) have oxidation number of **+1**

Metals of group **IIA** (Alkaline earth metals) have oxidation number of **+2**

Metals of group **IIIA** have oxidation number of **+3**

Oxidation number of metals of group **IVA**, Sn = +2, +4, Pb = +2, +4

Oxidation number of metal of group **VA**, Bi = +2

b. Generally, **metallic elements of group B** has oxidation number more than one type.
Example:

Table 8.1.
Oxidation numbers of several elements of group B

Elements of group B		*Oxidation numbers*
Name	*Symbol*	
Zink	*Zn*	+2
Silver	*Ag*	+1
Copper	*Cu*	+1, +2
Gold	*Au*	+1, +3
Iron	*Fe*	+2, +3
Lead	*Pb*	+2, +4

6. *Oxidation number of monoatomic ion*

Oxidation number of mono atomic ions is **equal to the charge on that ion**
Example:
Na^{+} ion has oxidation number of **+1**
Ba^{2+} ion has oxidation number of **+2**
Fe^{3+} ion has oxidation number of **+3**
Cl^{-} ion has oxidation number of –**1**
S^{2-} ion has oxidation number of –**2**

7. **The sum** of oxidation number of element atoms in **a compound molecule** is **equal to 0 (zero)**

$$\sum \text{o. n. of element in compound molecule} = 0$$

Example: H_2O

$$\begin{aligned}(\text{o.n. of H} \times 2) + (\text{o.n. of O} \times 1) &= 0\\ \{(+1) \times 2\} + \{(-2) \times 1\} &= 0\\ \{+2\} + \{-2\} &= 0\end{aligned}$$

8. **The sum** of oxidation number of element atoms in **a polyatomic ion** is **equal to the charge** on that ion.

$$\sum \text{o. n. of element in ion} = \text{charge of ion}$$

Example: OH^-

$$\begin{aligned}(\text{o.n. of O} \times 1) + (\text{o.n. of H} \times 1) &= -1\\ \{(-2) \times 1\} + \{(+1) \times 1\} &= -1\\ \{-2\} + \{+1\} &= -1\end{aligned}$$

The oxidation number of simple ion is the charge on that ion

E.g.

Cl^- is -1, Na^+ is +1

O^{2-} is -2 Ca^{2+} is +2

Al^{3+} is +3 N^{3-} is -3

S^{2-} is -2

The sum of oxidation numbers in a complex ion is the charge on the ion

SO_4^{2-} = -2 [total]

PO_4^{3-} = -3

NH_4^+ = +1

NO_3^- = -1

- Oxidation is a decrease in oxidation number
- Reduction is a gain of oxidation number

$$Mg + Cu^{2+} = Mg^{2+} + Cu$$

$$0 \qquad +2 \qquad +2 \qquad 0$$

- Mg has been oxidised Ox. No. zero to +2
- Loss of 2 electrons
- Cu has been reduced Ox. No. +2 to zero
- Gain of 2 electrons

Determining Oxidation Numbers of Elements

The oxidation number of an element in the molecule or in the ion, by use the rules of oxidation numbers can be determined.

•Write down the molecular or ionic formula which will be determined oxidation number its element and between one atom of element and the others, given enough space.

•Write each oxidation number of elements in below it and write x for element that will be determined its oxidation number.

•Use the rules of oxidation number, that is rule of number 7 or 8, for determine x value.

Example:

Determine the following element oxidation number

a. S in molecule of H_2SO_4

b. Cr in ion of $Cr_2O_7^{2-}$

Given : Molecule of H_2SO_4

Ion of $Cr_2O_7^{2-}$

Find : a. oxidation number of S in H_2SO_4

b. oxidation number of Cr in $Cr_2O_7^{2-}$

Solution :

a. H_2SO_4

o. n. H = +1, o. n. O = –2, o. n. S = x

$$\underset{+1}{H_2}\quad \underset{x}{S}\quad \underset{-2}{O_4}$$

$\sum$ o. n. element in molecule = 0

(2 x o. n. H) + (1 x o.n. S) + (4 x o.n. O) = 0

{ 2 x (+1) } + { 1 x (x) } + { 4 x (–2) = 0

(+2) + (x) + (–8) = 0

x = +8 – 2 ⇨ x = +6

⇨ The oxidation number of S in H_2SO_4 is **+6**

b. $Cr_2O_7^{2-}$

o. n. O = –2, o. n. Cr = x

$$(\underset{x}{Cr_2}\quad \underset{-2}{O_7})^{2-}$$

$\sum$ o. n. of element in ion = charge of ion

(2 x o. n. Cr) + (7 x o.n. O) = –2

{ 2 x (x) } + { 7 x (–2) } = –2

(2x) + (–14) = –2 ⇨ 2x = +14 - 2 ⇨ $x = \frac{+12}{2}$

x = + 6

⇨ The oxidation number of Cr in CrO_4^{2-} is **+6**

Determine the oxidation and reduction process , oxidising and reducing agent that occurs in the reactions below .

(a) $Mg + PbO \longrightarrow MgO + Pb$

(b) $Cu \longrightarrow Cu^{2+} + 2e^-$; $Cu^{2+} + 2e^- \longrightarrow Cu$

(c) $2CuO + C \longrightarrow 2Cu + CO_2$

(d) $Fe_2O_3 + 3C \longrightarrow 2Fe + 3CO$

(e) $Mg + CuO \longrightarrow MgO + Cu$

Test Yourself :

Calculate the oxidation number of the following elements :

(a)Manganese , Mn in potassium manganate (VII) , $KMnO_4$

(b)Manganese, Mn in manganate(VII) ion, MnO_4^{-}

(c)Chromium, Cr, in potassium dichromate(VI), $K_2Cr_2O_7$

(d)Cromium, Cr,in chromate(VI) ion, CrO_4^{2-}

(e)Iron in iron(II) chloride , $FeCl_2$

(f) Iron in iron(III) chloride , $FeCl_3$

(g)Carbon , C in sodium carbonate, Na_2CO_3

In each of the cases above, the oxidation number of each element is represented by the value of X .

Redox Reaction

In the chemical reaction, oxidation reaction and reduction reaction always occur together, it is called **oxidation reduction reaction** abreviated **as redox reaction.**

In the redox reaction occurs transfering of electrons from the substance that undergo oxidation to the substance that undergo reduction. Therefore, **redox reaction** is also called ***reaction of transfering electrons***

Special charateristic redox reaxtion is **the oxidation number change**.

- **Oxidation** : lossing electron, increasing oxidation number.
- **Reduction** : gaining electron, decreasing oxidation number.

The chemical reaction that does not espoused oxidation number change (increasing or decreasing in oxidation number) called **non-redox reaction.**

Example:

1. Redox reaction

Reaction of copper(II) oxide with hydrogen gas to form copper and water vapor

$$\underline{\mathbf{Cu}}O_{(s)} + \underline{\mathbf{H}}_{2(g)} \longrightarrow \underline{\mathbf{Cu}}_{(s)} + \underline{\mathbf{H}}_2O_{(g)} \quad \text{(redox)}$$

+2 0 0 +1

(***red***)

o. n. of Cu decreases from +2 to 0

***Changing** in o.n. of Cu is* **–2**

(***ox***)

o. n. of H increases from 0 to +1

***Total changing** in o.n. of H is* **+2**

➪ **In the redox reaction:**

total number of increasing in oxidation number in oxidation reaction = total number of decreasing in oxidation number in reduction reaction.

2. Non redox reaction

$$CuO_{(s)} + HCl_{(g)} \longrightarrow CuCl_{2(aq)} + H_2O_{(l)} \quad \text{(non-redox)}$$

$$Pb(NO_3)_{2(s)} + KI_{(g)} \longrightarrow PbI_{2(s)} + KNO_{3(aq)} \quad \text{(non-redox)}$$

Note : In the non-redox reaction above, **there are no oxidation number changes** of elements.

The oxidation numbers of elements in its compound is **constant**.

Oxidizing and Reducing Agents

If an element's oxidation number has increased in a reaction, this means the element has lost electrons and has been ***oxidized.***

If an element's oxidation number has decreased, this means it has gained electrons in the reaction and has been ***reduced.***

Example: $2Mg + O_2 \longrightarrow 2MgO$

$2Mg^0 \longrightarrow 2Mg^{2+}$ Mg has been oxidized

$O_2^0 \longrightarrow 2O^{2-}$ O has been reduced

Oxidizing and Reducing Agents

Example: Identify the oxidizing agents and reducing agents in the following reactions:

1. $Zn\ (s) + CuSO_4\ (aq) \longrightarrow ZnSO_4\ (aq) + Cu\ (s)$

$Zn \longrightarrow Zn^{2+}$ — Zn is oxidized — Zn is the ***reducing agent***

$Cu^{2+} \longrightarrow Cu$ — Cu^{2+} is reduced — $CuSO_4$ is the ***oxidizing agent***

2. $2\ K\ (s) + 2\ H_2O\ (aq) \longrightarrow 2\ KOH\ (aq) + H_2\ (g)$

$K \longrightarrow K^{+}$ — K is oxidized — K is the **reducing agent**

$H^{+} \longrightarrow H^{o}$ — H^{+} is reduced — H_2O is the **oxidizing agent**

List of common Oxidising and Reducing Agents

Oxidising agents	Reducing agents
Oxygen (O_2)	Hydrogen (H_2)
Chlorine ($Cl_2 \rightarrow Cl^-$)	Reactive metals ($M \rightarrow M^{n+}$)
Dichromate(VI) ions in acid ($Cr_2O_7^{2-} \rightarrow Cr^{3+}$)	Carbon monoxide ($CO \rightarrow CO_2$)
Manganate(VII) ions in acid ($MnO_4^- \rightarrow Mn^{2+}$)	Carbon (C)
Manganate(VII) ions in neutral or alkali ($MnO_4^- \rightarrow MnO_2$)	Iron(II) salts ($Fe^{2+} \rightarrow Fe^{3+}$)
Manganese(IV) oxide in acid ($MnO_2 \rightarrow Mn^{2+}$)	Thiosulphate ions to tetrathionate ions ($S_2O_3^{2-} \rightarrow S_4O_6^{2-}$)
Hydrogen peroxide ($H_2O_2 \rightarrow H_2O$)	Ethanedioate ions ($C_2O_4^{2-} \rightarrow CO_2$)
Concentrated sulphuric acid ($H_2SO_4 \rightarrow SO_2$)	Hydrogen peroxide ($H_2O_2 \rightarrow O_2$)
Iron(III) salts ($Fe^{3+} \rightarrow Fe^{2+}$)	Sulphate(IV) to sulphate(VI) ($SO_3^{2-} \rightarrow SO_4^{2-}$)
	Iodide to iodine ($I^- \rightarrow I_2$)

Realise something?

H_2O_2 is both an oxidising and a reducing agent!

If a stronger oxidising agent is present, H_2O_2 is reducing

Oxidizing and Reducing Agents

The element that was oxidized "donated" its electrons to the element that was reduced (gained electrons). Thus, the species that contains the element being oxidized is said to be the reducing agent.

The element that was reduced "stole" electrons from the element that was oxidized (lost electrons). Thus, the species that contains the element being reduced is said to be the oxidizing agent.

The more readily an element is oxidized, the better it is as a reducing agent. Conversely, the more readily an element is reduced, the better it is as an oxidizing agent.

Oxidizing Agent (Oxidant) and Reducing Agent (Reductant)

The reactants that involve in a redox reaction can be differentiated into two kinds, that is oxidizing agent (oxidant) and reducing agent (reductant)

Oxidizing agent (oxidant)

Oxidizing agent is:

- a reactant that **oxidizes** other reactant
- a reactant that can **gain electron**
- a reactant that in a reaction undergoes **reduction**
- a reactant that in a reaction undergoes **decreasing** in **oxidation number**

Examples:

Halogen, F_2, Cl_2, Br_2, I_2

Oxygen, O_2

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

$$\underbrace{0 \qquad\qquad\qquad\qquad\qquad (-1) \times 2}_{\text{(o. n. of Cl decreases 0 to -1)}}$$

(reduction)

Cl_2 is **oxidizing agent** (**oxidant**),
because in that reaction Cl_2 undergoes **reduction** or **decreasing** in **oxidation number**, from 0 to -1

•Reducing agent (reductant)

Reducing agent is:

- ❖ a substance (reactant) that **reduces** other substances (reactants)
- ❖ a substance (reactant) that can **loss electron**
- ❖ a substance (reactant) that in the reaction undergoes **oxidation**
- ❖ a substance (reactant) that undergoes **increasing** in **oxidation number**

Example:

Hydrogen, H_2

Ion halides; F^-, Cl^-, Br^-, I^-

metals

$$H_{2(g)} \longrightarrow 2\,H^+_{(aq)} + 2\,e^-$$

$$0 \qquad\qquad (+1) \times 2$$

(oxidation)

o. n. of H increases from 0 to +1

H_2 is **reducing agent** (**reductant**),
because in that reaction H_2 undergoes **oxidation** or **increasing** in **oxidation number,** from 0 to +1

Example problem :

Given a redox reaction:

$$3S_{(s)} + 2KClO_{3(s)} \longrightarrow 3SO_{2(g)} + 2KCl_{(s)}$$

a. Identify and under line, element atoms of reactants undergo change in oxidation number.

b. Determine the reactants that undergo reduction - oxidation include their product, and calculate its oxidation number change

c. Determine the reactant behaves as oxidant and reductant.

Answer:

a. In the redox reaction:

$$\underset{0}{3\,\underline{\mathbf{S}}_{(s)}} + \underset{(+5)}{2\,K\underline{\mathbf{Cl}}O_{3(s)}} \longrightarrow \underset{(+4)}{3\,\underline{\mathbf{S}}O_{2(g)}} + \underset{(-1)}{2\,K\underline{\mathbf{Cl}}_{(s)}}$$

Element atoms undergo change in oxidation number is:

- S : oxidation number of S **increases** from 0 to +4
- Cl : oxidation number of Cl element atom in $KClO_3$ **decreases** from +5 to -1

b. In the redox reaction:

$$3\,\underline{S}_{(s)} + 2\,K\underline{Cl}O_{3(s)} \longrightarrow 3\,\underline{S}O_{2(g)} + 2\,K\underline{Cl}_{(s)}$$

0 (+5) (+4) (-1)

(0) (Ox) S is oxidized into SO_2 (+12)

The total increasing o.n. of S (three atoms) is +12

(+10) $KClO_3$ is reduced into KCl (Red) (-2)

The total decreasing o.n. of Cl (two atoms) is -12

c. In the redox reaction:

$$3\,\underline{S}_{(s)} + 2\,K\underline{Cl}O_{3(s)} \longrightarrow 3\,\underline{S}O_{2(g)} + 2\,K\underline{Cl}_{(s)}$$

0 (+5) (+4) (-1)

S undergoes **oxidation**

The element of **S** is **reducing agent**

$KClO_3$ undergoes **reduction**

The compound of **$KClO_3$** is **oxidizing agent**

Auto Redox Reaction (Disproportionation)

Auto redox reaction is a reaction of reduction and oxidation that occur in the same substance (reactant).

Example of auto redox reaction:

Reaction of chlorine gas with sodium hydroxide solution

$$\underline{\mathbf{Cl}}_{2(g)} + 2\,NaOH_{(aq)} \longrightarrow Na\underline{\mathbf{Cl}}_{(aq)} + Na\underline{\mathbf{Cl}}O_{(aq)} + H_2O_{(l)}$$

$0 \qquad\qquad -1 \qquad\qquad +1$

(reduction)

o. n. of Cl ***decreases*** *from 0 into –1*

(oxidation)

o. n. of Cl ***increases*** *from 0 into + 1*

IUPAC Nomenclature

The compound that is formed by the elements have more than one type of oxidation number , its name diferentiated by the Roman number writing in the bracket in the back of that element name. The **Roman number** shows the **value of oxidation number** of that element.

The compound that is formed by the element only has one type of oxidation number, the Roman number does not need writen.

This **IUPAC nomenclature** applies in both **ionic and covalent compounds.**

Examples **IUPAC name** of **binary covalent compound**:

CO	: carbon(II) oxide	(oxidation number of C = +2)
CO_2	: carbon(IV) oxide	(oxidation number of C = +4)
P_2O_3	: phosphorus(III) oxide	(oxidation number of P = +3)
N_2O_5	: nitrogen(V) oxide	(oxidation number of N = +5)
Cl_2O_7	: chlorine(VII) oxide	(oxidation number of Cl = +7)

Examples **IUPAC name** of **binary ionic compounds**

$ZnCl_2$ = zink chloride

Al_2O_3 = aluminium oxide

Cu_2O = copper(I) oxide

CuS = copper(II) sulfide

Identify and under line, **element atoms** of substaces of the following reaction undergo **change in oxidation number.**

$$CrI_{3(ag)} + KOH_{(aq)} + Cl_{2(g)} \longrightarrow KCrO_{4(aq)} + KIO_{4(aq)} + KCl_{4(aq)} + H_2O_{(l)}$$

Example: Consider the compound PCl_3

Cl is the more electronegative element, so assign each Cl an oxidation # equal to its charge as an anion (= –1).

PCl_3 is a neutral molecule, so the sum of the oxidation numbers for P and Cl must add to 0.

Let x = oxidation # of P: $x + 3(-1) = 0$ ⑧ $x = +3$

$\boxed{PCl_3: \quad P = +3, \; Cl = -1}$

Redox Reactions

example: Consider $HClO_4$.

Since oxygen is present, but not as a peroxide ion, the oxygen is assigned an oxidation # of –2. The H is not bonded to a metal, so it must have an oxidation # of +1. Let x = oxidation number of Cl, and note that the sum of the oxidation #'s must equal zero since $HClO_4$ has a *net* charge of zero:

$0 = +1 + x + 4(-2)$ solving for x gives $x = +7$ = Cl.

$HClO_4$: H = +1, Cl = +7, O = -2

Note carefully that Cl is NOT a +7 cation in $HClO_4$!! The +7 oxidation state simply tells us the electron density around Cl in $HClO_4$ is significantly lower than the electron density around Cl in its elemental state.

Determine the oxidation state of…

1) H in H_2O
2) N in NH_4^+
3) S in $S_2O_3^{2-}$
4) Cr in $Cr_2O_7^{2-}$

Determine the oxidation state of...

* Sum of all oxidation states of all atoms = Overall Charge of molecule / ion / atom

1) H in H_2O

Let the oxidation state of H be x.

Thus, in H_2O, $2x + (-2) = 0$

$x = 1$

Determine the oxidation state of…

* Sum of all oxidation states of all atoms = Overall Charge of molecule / ion / atom

2) N in NH_4^+

Let the oxidation state of N be *x*.

Thus, in NH_4^+, $x + 4(+1) = +1$

$x = -3$

Determine the oxidation state of…

* **Sum of all oxidation states of all atoms = Overall Charge of molecule / ion / atom**

3) S in $S_2O_3^{2-}$

Let the oxidation state of S be *x*.

Thus, in $S_2O_3^{2-}$, $2x + 3(-2) = -2$

$x = +2$

Determine the oxidation state of...

* **Sum of all oxidation states of all atoms = Overall Charge of molecule / ion / atom**

4) Cr in $Cr_2O_7^{2-}$

Let the oxidation state of Cr be x.

Thus, in $Cr_2O_7^{2-}$, $2x + 7(-2) = -2$

$x = +6$

Example 1

$$MnO_4^{-} + 5Fe^{2+} + 8H^{+} \longrightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$$

+7   +2   +2   +3

Let the oxidation state of Mn be *x*.

Thus, in MnO_4^-, *x* + 4(-2) = -1

***x* = +7**

- **<u>Manganese is reduced from oxidation state of +7 in MnO_4^- to +2 in Mn^{2+}, while iron is oxidised from oxidation state of +2 in Fe^{2+} to +3 in Fe^{3+}.</u>**

Special Redox: Disproportionation

Definition:

A disproportionation reaction is a redox reaction in which one species is simultaneously oxidised and reduced.

$$Cl_2 + 2OH^- \longrightarrow ClO^- + Cl^- + H_2O$$

0 **+1** **-1**

Chlorine is simultaneously reduced from oxidation state of 0 in Cl_2 to -1 in Cl^-, and oxidised from oxidation state of 0 in Cl_2 to +1 in ClO^-.

Balancing redox reactions

Example:
Try to balance the following reaction by trial and error.

$$MnO_4^{-} + H_2O_2 + H^{+} \longrightarrow Mn^{2+} + O_2 + H_2O$$

Possible answer:

$$MnO_4^{-} + H_2O_2 + 2H^{+} \rightarrow Mn^{2+} + 2O_2 + 2H_2O$$

$$2MnO_4^{-} + 4H_2O_2 + 4H^{+} \rightarrow 2Mn^{2+} + 3O_2 + 6H_2O$$

Balancing redox reactions

- **Example:**
  Try to balance the following reaction by trial and error.

$$MnO_4^- + H_2O_2 + H^+ \longrightarrow Mn^{2+} + O_2 + H_2O$$

- **Actual answer:**

$$2MnO_4^- + 5H_2O_2 + 6H^+ \longrightarrow 2Mn^{2+} + 5O_2 + 8H_2O$$

- **Note: You might not even be told at the beginning that H^+ is reactant, H_2O is product.**

The half-equation method

- Write down the given reactants and products of the reaction
- Identify the atoms in the given species that are undergoing oxidation / reduction and construct the unbalanced oxidation / reduction half-equations
- Balance both the half-equations using the following steps:
  - Balance the "odd" atoms ("odd" atoms refer to atoms other than oxygen and hydrogen)
  - Balance oxygen atoms by adding H_2O molecules
  - Balance hydrogen atoms by adding H^+ ions
  - Balance charges by adding electrons
- Multiply the balanced half-equations by appropriate integers such that the number of electrons in both half-equations are equal
- Add the resulting half-equations together, and eliminate any common species on both sides to obtain the balanced equation.

The half-equation method

- <u>Example:</u>
- **Balance the following reaction:**

$$MnO_4^- + H_2O_2 \longrightarrow Mn^{2+} + O_2$$

The half-equation method

- **Step 1:** Write down the given reactants and products of the reaction

$$MnO_4^{-} + H_2O_2 \longrightarrow Mn^{2+} + O_2$$

The half-equation method

- Step 2: **Identify the atoms in the given species that are undergoing oxidation / reduction and write the unbalanced oxidation / reduction half-equations**

$$MnO_4^{-} + H_2O_2 \longrightarrow Mn^{2+} + O_2$$

Reduced

Oxidised

- **Reduction half-equation:** $MnO_4^{-} \longrightarrow Mn^{2+}$

- **Oxidation half-equation**: $H_2O_2 \longrightarrow O_2$

The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance the atoms undergoing oxidation / reduction

- **Reduction half-equation**:

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

The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance oxygen atoms by adding H_2O molecules

- **Reduction half-equation:**

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

The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance hydrogen atoms by adding H^+ ions

- **Reduction half-equation:**

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

The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance hydrogen atoms by adding H^+ ions

- **Reduction half-equation:**

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

The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance charges by adding electrons

- **Reduction half-equation**:

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

The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance charges by adding electrons

- **Reduction half-equation**:

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

The half-equation method

- Step 3: Balance both the half-equations using the following steps:

- **Reduction half-equation**:

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

- **Oxidation half-equation:**

$$H_2O_2 \longrightarrow O_2$$

The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance the atoms undergoing oxidation / reduction
- **Reduction half-equation**:

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

- **Oxidation half-equation**:

$$H_2O_2 \longrightarrow O_2$$

The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance oxygen atoms by adding H_2O molecules
- **Reduction half-equation:**

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

- **Oxidation half-equation**:

$$H_2O_2 \longrightarrow O_2$$

The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance hydrogen atoms by adding H^+ ions

- **Reduction half-equation**:

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

- **Oxidation half-equation**:

$$H_2O_2 \longrightarrow O_2$$

The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance hydrogen atoms by adding H^+ ions

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

- **Oxidation half-equation**:

$$H_2O_2 \longrightarrow O_2 + 2H^+$$

The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance charges by adding electrons

- **Reduction half-equation:**

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

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The half-equation method

- Step 3: Balance both the half-equations using the following steps:
  - Balance charges by adding electrons

- **Reduction half-equation:**

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

- **Oxidation half-equation:**

$$H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-$$

The half-equation method

- Step 4: Multiply the balanced half-equations by appropriate integers such that the number of electrons in both half-equations are equal

- **Reduction half-equation:**

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

- **Oxidation half-equation**:

$$H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-$$

The half-equation method

- Step 4: Multiply the balanced half-equations by appropriate integers such that the number of electrons in both half-equations are equal

- **Reduction half-equation:**

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

- **Oxidation half-equation:**

$$H_2O_2 \longrightarrow O_2 + 2H^{+} + 2e^{-}$$

The half-equation method

- Step 4: Multiply the balanced half-equations by appropriate integers such that the number of electrons in both half-equations are equal

- **Reduction half-equation:**

$$\left[MnO_4^{\ -} + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O\right] \text{ x 2}$$

- **Oxidation half-equation:**

$$\left[H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-\right] \text{ x 5}$$

The half-equation method

- Step 4: Multiply the balanced half-equations by appropriate integers such that the number of electrons in both half-equations are equal

- Reduction half-equation:

x 2

$$2MnO_4^- + 16H^+ + 10e^- \longrightarrow 2Mn^{2+} + 8H_2O$$

- **Oxidation half-equation:**

$$H_2O_2 \longrightarrow \left(O_2 + 2H^+ + 2e^- \right)$$

x 5

The half-equation method

- Step 4: Multiply the balanced half-equations by appropriate integers such that the number of electrons in both half-equations are equal

- **Reduction half-equation:** x 2

$$2MnO_4^- + 16H^+ + 10e^- \longrightarrow 2Mn^{2+} + 8H_2O$$

- **Oxidation half-equation**: x 5

$$5H_2O_2 \longrightarrow 5O_2 + 10H^+ + 10e^-$$

The half-equation method

- Step 5: Add the resulting half-equations together, and eliminate any common species on both sides to obtain the balanced equation.

- **Reduction half-equation:** x 2

$$2MnO_4^- + 16H^+ + 10e^- \longrightarrow 2Mn^{2+} + 8H_2O$$

- **Oxidation half-equation:** x 5

$$5H_2O_2 \longrightarrow 5O_2 + 10H^+ + 10e^-$$

The half-equation method

- Step 5: Add the resulting half-equations together, and eliminate any common species on both sides to obtain the balanced equation.

$$2MnO_4^- + 6H^+ + \cancel{10e^-} + 5H_2O_2 \longrightarrow 2Mn^{2+} + 8H_2O + 5O_2 + \cancel{10H^+} + \cancel{10e^-}$$

The half-equation method

- Step 5: Add the resulting half-equations together, and eliminate any common species on both sides to obtain the balanced equation.

- **Balanced Equation:**

$$2MnO_4^- + 6H^+ + \cancel{10e^-} + 5H_2O_2 \longrightarrow 2Mn^{2+} + 8H_2O + 5O_2 + \cancel{10H^+} + \cancel{10e^-}$$

The half-equation method

- Step 5: Add the resulting half-equations together, and eliminate any common species on both sides to obtain the balanced equation.

- **Balanced Equation:**

$$2MnO_4^- + 5H_2O_2 + 6H^+ \longrightarrow 2Mn^{2+} + 5O_2 + 8H_2O$$