CHEMISTRY ONLINE
— TUITION —

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# CHEMISTRY

## REVISION NOTES

## REDOX

# Redox

## Oxidation:

- Oxidation is the process in which an atom, ion, or molecule loses electrons.
- It results in an increase in the oxidation state of the species involved.
- Oxidation often involves the addition of oxygen or loss of hydrogen.

## Reduction:

- Reduction is the process in which an atom, ion, or molecule gains electrons.
- It results in a decrease in the oxidation state of the species involved.
- Reduction often involves the addition of hydrogen or loss of oxygen.

## Oxidizing Agent

An oxidising agent is a substance that accepts electrons from another substance during a chemical reaction.

## Reducing Agent

Now, this one is the electron donor. A reducing agent is a substance that gives away electrons to another substance during a chemical reaction.

So, in terms of electrons:

- **Oxidising Agent:** Gains electrons (by making others lose electrons).
- **Reducing Agent:** Loses electrons (by giving them away to others)

## Rules for Assigning Oxidation States:

1. **Elements in their natural state:** In their elemental form, atoms have an oxidation state of 0. For example, $O_2$, $H_2$, $N_2$, $Cl_2$ all have oxidation states of 0.

2. **Monoatomic ions:** The oxidation state of a monatomic ion is equal to its charge. For example, the oxidation state of $Na^+$ is +1, and the oxidation state of $Cl^-$ is -1.

3. **Hydrogen:** Hydrogen typically has an oxidation state of +1 when combined with nonmetals and -1 when combined with metals.

4. **Oxygen:** Oxygen typically has an oxidation state of -2 in compounds. There are some exceptions, such as in peroxides (e.g., $H_2O_2$), where its oxidation state is -1.

5. **Alkali metals and alkaline earth metals:** Alkali metals (e.g., Li, Na, K) have an oxidation state of +1, and alkaline earth metals (e.g., Mg, Ca) have an oxidation state of +2.

6. **Fluorine:** Fluorine always has an oxidation state of -1 in compounds.

7. **The sum of oxidation states:** In a neutral compound, the sum of the oxidation states of all atoms must equal zero. In a polyatomic ion, the sum of oxidation states should equal the charge of the ion.

8. **Oxidation states in complex ions:** In complex ions, consider the charge on the ion as a whole when determining the oxidation state of each element within the ion.

9. **Redox reactions:** In a redox reaction, the substance that is oxidised has its oxidation state increased, while the substance that is reduced has its oxidation state decreased.

10. **Change in oxidation state:** Be aware of the change in oxidation state for each element in a chemical reaction. The change in oxidation state is equal to the number of electrons transferred.

## Redox Equation (Reduction-Oxidation Equation):

A redox equation represents a chemical reaction in which there is a transfer of electrons from one substance (reductant or reducing agent) to another substance (oxidant or oxidizing agent).

In a redox equation, you typically have two half-reactions: one representing the oxidation half (loss of electrons), and the other representing the reduction half (gain of electrons).

These half-reactions are balanced so that the number of electrons lost in the oxidation half is equal to the number of electrons gained in the reduction half.

For example:

**Half-Reaction 1 (Oxidation):**

$Cu \rightarrow Cu^{2+} + 2e^-$

**Half-Reaction 2 (Reduction):**

$Ag^+ + e^- \rightarrow Ag$

**Overall Redox Equation:**

$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$

This is a simple example of a redox equation involving the oxidation of copper (Cu) and the reduction of silver ions ($Ag^+$).

## Balancing more complex equations.

**Balanced Reduction Half-Equation for $NO_3^-$ to $NO_2^-$**

1. Write the initial half-equation without balancing electrons

 

$NO_3^- \rightarrow NO_2^-$

2. Balance the nitrogen atoms by adding a coefficient of 1 to NO3- and NO2-:

$NO_3- \rightarrow NO_2-$

3. Balance the oxygen atoms by adding water (H2O) to the product side:

$NO_3- \rightarrow NO_2- + H_2O$

4. Balance the hydrogen atoms by adding hydrogen ions (H+) to the reactant side:

$NO_3- + H+ \rightarrow NO_2- + H_2O$

5. Now, add electrons (e-) to balance the change in oxidation number, which is 2 electrons in this case.

   $NO_3- + H+ + 2e- \rightarrow NO_2- + H_2O$

   This balanced half-equation represents the reduction of nitrate ion ($NO_3$-) to nitrite ion ($NO_2$-) by first balancing the atoms and then adding 2 electrons to balance the change in oxidation number.

**<u>Oxidation Half-Equation for Zn to $Zn^{2+}$</u>**

1. Identify the oxidation numbers and changes:

   Zinc (Zn) goes from 0 to +2 (an oxidation).

2. Write the initial half-equation without balancing electrons:

   $Zn \rightarrow Zn^{2+}$

3. Add electrons (e-) to balance the change in oxidation number:

   $Zn \rightarrow Zn^{2+} + 2e-$

4. This half-equation represents the oxidation of zinc (Zn) to form zinc ions (Zn^2+) with the loss of 2 electrons

**<u>Combine oxidation and reduction half-equations for zinc (Zn) and nitrate ($NO^{3-}$) to nitrite ($NO^{2-}$) into a complete redox equation.</u>**

Here are the balanced half-equations for reference:

**Oxidation Half-Equation for Zn to $Zn^{2+:}$**

- $Zn \rightarrow Zn^{2+} + 2e-$

**Reduction Half-Equation for $NO^{3-}$ to NO2-:**

$NO^{3-} + H+ + 2e- \rightarrow NO^{2}- + H_2O$

 

To combine them, you must ensure that the number of electrons lost in the oxidation half-equation matches the number of electrons gained in the reduction half-equation.

You can do this by multiplying one or both of the equations to ensure the electrons are equal.

Since the oxidation half-equation involves the loss of 2 electrons, and the reduction half-equation gains 2 electrons, you can combine them directly without any need to multiply:

$Zn + NO^{3-} + H+ \rightarrow Zn^{2+} + NO_2- + H_2O$

This is the balanced redox equation for the reaction in which zinc (Zn) is oxidised to form zinc ions ($Zn^2+$), and nitrate ions ($NO^{3-}$) are reduced to form nitrite ions ($NO^{2-}$) in the presence of hydrogen ions (H+) and water ($H_2O$).

## Disproportionation Reaction

A disproportionation reaction is a type of redox (reduction-oxidation) reaction in which a single substance undergoes both oxidation and reduction, leading to the formation of two different products with distinct oxidation states of the same element.

**Example 1:**

Consider the disproportionation of hydrogen peroxide ($2H_2O_2$)

$$2H_2O_2 \longrightarrow 2H_2O + O_2$$

In this reaction, hydrogen peroxide ($2H_2O_2$ acts as both the oxidising and reducing agent. One oxygen atom in $2H_2O_2$ is oxidised (loses electrons) to form water ($H_2O$), while the other oxygen atom is reduced (gains electrons) to form molecular oxygen.The net result is the disproportionation of hydrogen peroxide into water and oxygen.

**Example 2:**

Disproportionation reaction involving chlorine and water:

$Cl2+H2O \longrightarrow HCl+HOCl$

In this reaction, chlorine ($2Cl2$) undergoes disproportionation. One chlorine atom is reduced (gains electrons) to form hydrochloric acid (*HCl*), while the other chlorine atom is oxidized (loses electrons) to form hypochlorous acid (*HOCl*). The net result is the formation of two different products with distinct oxidation states of chlorine, illustrating a disproportionation reaction.

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## Exam Questions

State what happens to a reducing agent during a reaction, in terms of oxidation number **and** electrons.

(1)

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Sodium chlorate(I) is a bleaching agent.

(i) Sodium chlorate(I) can be made by the reaction of chlorine with sodium hydroxide.

Show, by using oxidation numbers, that this reaction is disproportionation.

$$2NaOH + Cl_2 \rightarrow NaClO + NaCl + H_2O$$

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(i) In a dry container, a fluoride of silver reacts with sulfur to produce disulfur difluoride. Complete the equation for this reaction. State symbols are not required.

(1)

$$S_8 + \ldots\ldots AgF_2 \rightarrow \ldots\ldots S_2F_2 + \ldots\ldots AgF$$

(ii) Explain, by using the oxidation numbers of **all** the atoms, whether or not this is a redox reaction.

(3)

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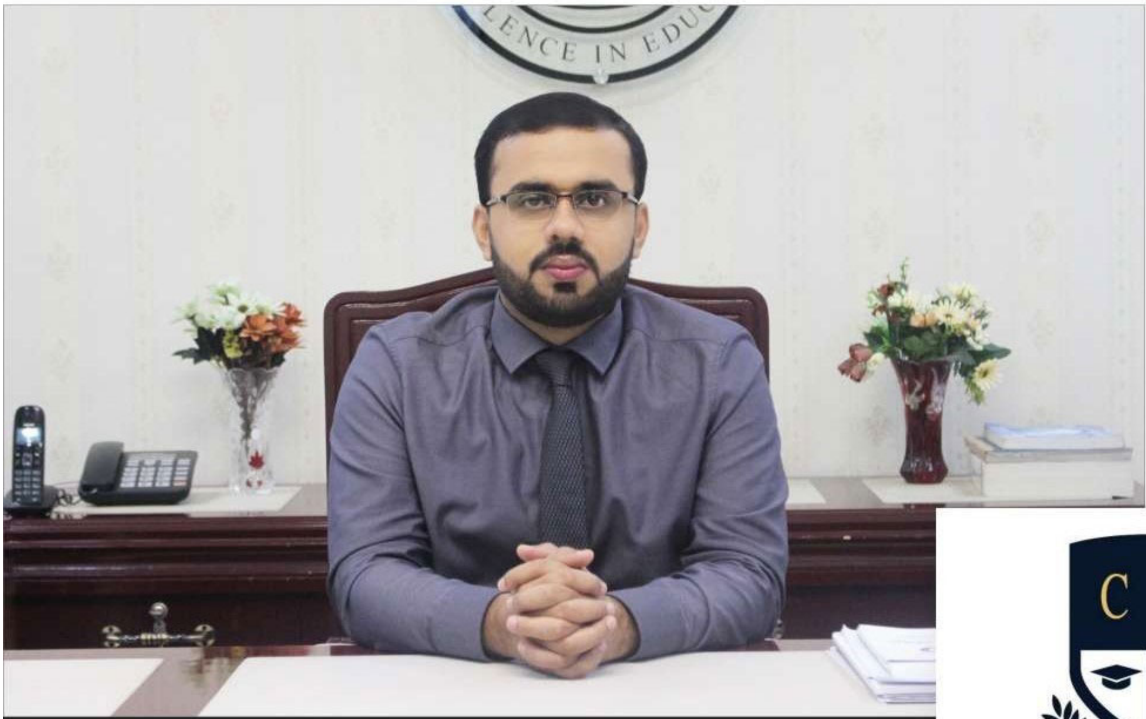
Calcium reacts with chlorine.

$$Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$$

Explain, in terms of electrons, why this is a redox reaction.



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