Redox Reactions

Definitions of Redox

Type	Definition of Oxidation	Definition of Reduction
Oxygen	Gain of oxygen	Loss of oxygen
Hydrogen	Loss of hydrogen	Gain of hydrogen
Electrons	Loss of electrons	Gain of electrons
Oxidation States	Increase in Oxidation State	Decrease in Oxidation State

Transfer of Electrons - Definition

In the following redox reaction, **zinc metal is oxidised while copper(II) ions are reduced.**

Consider the following reaction:

$$\ce{Zn(s) + CuCl_2(aq) -> ZnCl_2 (aq) + Cu(s)}$$

Ionic equation:

$$\ce{Zn(s) + Cu^{2+} (aq) -> Zn^{2+} (aq) + Cu(s)}$$

An ionic half equation shows which reactant particles gain or lose electrons to form products. To construct ionic half equations, follow these steps:

1. **Isolate reactant and product of the same element** into a single ionic half equation.
2. **Add coefficients** to balance the number of elements on the reactant and product sides of the ionic half equation
3. **Add electrons** to the ionic half equations to balance the charges on the reactant and product sides. Note that electrons are negatively charged.
4. **Write state symbols** for all reactant and product particles.

Oxidation ionic half equation involving zinc:

$$\ce{Zn(s) -> Zn^{2+} (aq) + 2e^-}$$

Reduction ionic half equation involving copper:

$$\ce{Cu^{2+} + 2e^- -> Cu(s)}$$

Reactant Particle	Is reactant particle oxidised/reduced?	Are electrons on reactant or product side?	Did reactant particle undergo gain or loss of electrons?
$\ce{Zn}$	Oxidised	product	loss
$\ce{Cu^{2+}}$	Reduced	reactant	gain

Using ideas about electron transfer, explain why the reaction between zinc metal and aqueous copper(II) chloride is a redox reaction.

- Zinc atom loses electrons to form $\ce{Zn^{2+}}$ ions and copper(II) ions gained electrons to form copper atoms. Hence, zinc atom is oxidised and copper(II) ions are reduced simultaneously.

Based on the definitions of oxidation and reduction written in (iii), deduce if oxidation and reduction can take place independently of one another. Explain your answer.

- Oxidation and reduction cannot take place independently of one another as electrons that are lost by a particle is gained by another particle.

From your answer in (v), briefly define a redox reaction.

- A redox reaction is a chemical reaction in which reduction and oxidation occur at the same time.

Examples

By constructing ionic and ionic half equations for the following reactions, use ideas about electron transfer to explain why they are redox reactions.

$$\ce{2HCl(aq) + Mg(s) -> MgCl_2(aq) + H_2(g)}$$

Ionic equation: $\ce{2H^+(aq) + Mg(s) -> Mg^{2+}(aq) + H_2(g)}$

Ionic half equations:

Oxidation: $\ce{Mg(s) -> Mg^{2+} (aq) + 2e^-}$

Reduction: $\ce{2H^+(aq) + 2e^- -> H_2(g)}$

Using ideas about electron transfer, explain why the reaction is a redox reaction.

Magnesium atom loses electrons and is oxidised to form $\ce{Mg^{2+}}$ ions, while hydrogen ions gains electrons and is reduced to $\ce{H_2}$ molecules simultaneously.

Oxidation State Rules

1. Elements that are uncombined with other elements are assigned an oxidation number of 0.
2. In simple ions, the oxidation number is simply the charge on the ion.
3. In complex ions, the sum of the oxidation numbers is equivalent to the net charge for polyatomic ions.
4. The oxidation number of hydrogen in all of its compounds is +1, except in metal hydrides where its oxidation number is -1.
5. The oxidation number of oxygen in all its compounds is -2, except in peroxides where it is -1.

6. The oxidation number of Group 1 elements (e.g. sodium) in their compounds is +1, for Group 2 elements (e.g. magnesium) in their compounds is +2, and for aluminium in its compounds is +3.
7. There are many oxidation numbers for Group 17 elements in their compounds but the usual one is -1.
8. The sum of the oxidation numbers of all the elements in a compound is zero.

Explaining Redox Using O.S

> The oxidation state of chlorine decreased from 0 in $\ce{Cl_2}$ to -1 in $\ce{KCl}$, hence chlorine was reduced. The oxidation state of bromine increased from -1 in $\ce{KBr}$ to 0 in $\ce{Br_2}$, hence bromine was oxidised simultaneously.

Oxidising and Reducing Agents

In a redox reaction, both oxidation and reduction take place.

- An **oxidising agent** (or oxidant) oxidises another substance - and is itself reduced.
- An **reducing agent** (or reductant) reduces another substance - and is itself oxidised.

Example

Identify the oxidising agent and reducing agent in the following reaction:

$$\ce{CuO (s) + H_2 (g) -> Cu (s) + H_2 O (\textit{l})}$$

The oxidation state of **hydrogen** increased from 0 in $\ce{H_2}$ to +1 in $\ce{H_2 O}$. Hence, hydrogen is oxidised. The oxidation state of **copper** decreased from +2 in $\ce{CuO}$ to 0 in $\ce{Cu}$. Therefore, hydrogen is the reducing agent and $\ce{CuO}$ is the oxidising agent.

Testing of Oxidising Agents and Reducing Agents

Acidified potassium manganate(VII) solution

Half equation: $\ce{MnO_4^- (aq) + 8H^+ (aq) + 5e^- -> Mn^{2+} (aq) + H_2 O (\textit{l})}$

Acidified $\ce{KMnO_4}$ solution is a strong oxidising agent and it turns from purple to colourless when coming into contact with a reducing agent.

Potassium iodide solution

Half equation: $\ce{2I^- (aq) -> I_2 (s) + 2e^-}$

$\ce{KI}$ solution is a reducing agent and it turns from colourless to brown when coming into contact with an oxidising agent.