Year 11
Chemistry
Reactive Chemistry

Lesson 5
Oxidation and reduction
Sample resources

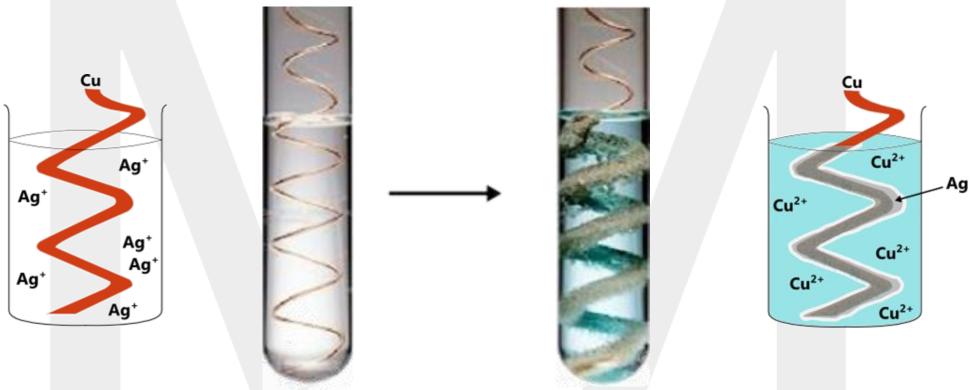
MATRIX
EDUCATION

1. Oxidation and reduction (redox) reactions

☐ Oxidation and reduction

- Recall that a more active metal will displace a less active metal from a compound.
  - For example, copper displaces silver from silver nitrate solution according to the following chemical equation:

$$Cu_{(s)} \quad + \quad 2AgNO_{3(aq)} \quad \longrightarrow \quad 2Ag_{(s)} \quad + \quad Cu(NO_3)_{2(aq)}$$

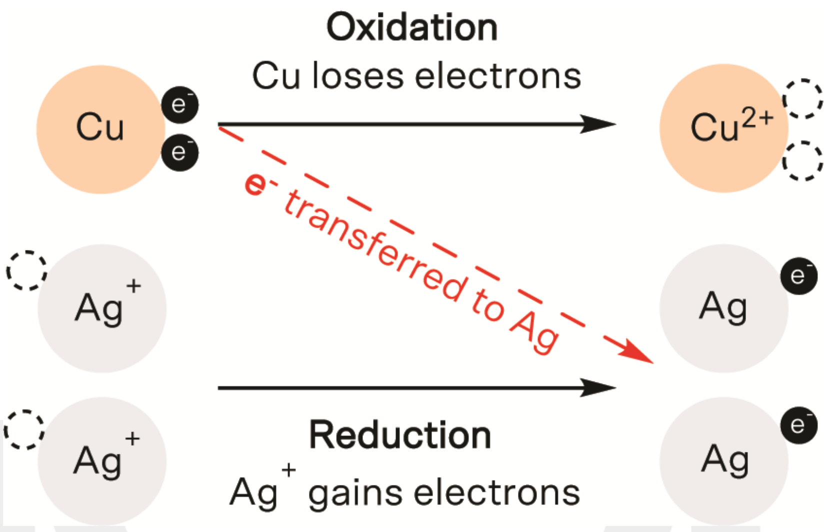


- Displacement reactions belong to a class of reactions known as **redox (reduction-oxidation) reactions.**
  - These reactions involve the **transfer of electrons.** One chemical species loses electrons, which are gained by another chemical species.
  - These processes occur simultaneously.
  - The other metal reactions (with water, acid and oxygen) are all redox reactions too.

- Species that **lose electrons** are **oxidised.**

- Species that **gain electrons** are **reduced.**

- The following diagram shows the redox reaction between copper and silver ions:



  - A copper atom **transfers two electrons** to two silver ions, producing a copper ion and two silver atoms.
  - The copper atom has **lost two electrons (oxidation)**.
  - Each silver ion has **gained one electron (reduction)**.

- In the reaction of copper and silver nitrate on the previous page:
  - Which substance has been oxidised?

  ____

  - Which substance has been reduced?

  ____

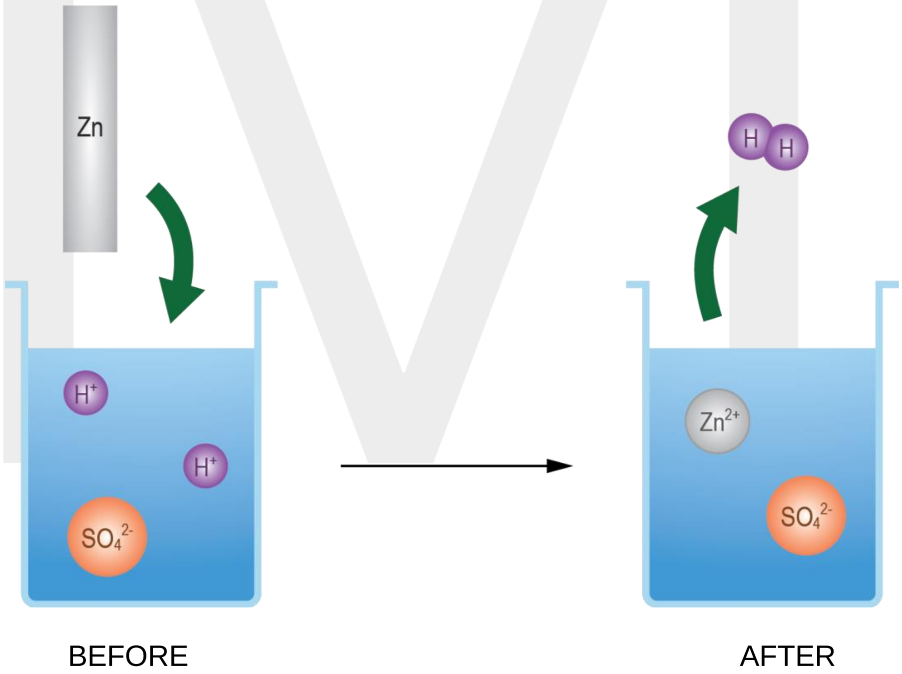
☐ Half-equations

- A redox process can be represented using **half-equations** (ion-electron equations) to show the **loss or gain of electrons**.
  - There will always be two half-equations in a redox reaction.
  - In a two-reactant reaction, there is typically one half-equation per reactant.

- Like net ionic equations, half-equations must be **balanced** as to both number of atoms and charge.
  - The net ionic equation can be **split up** to give the half-equations.
  - The half-equations can also be **combined** to give the **net ionic equation** which does not contain electrons.

- For example, when zinc reacts with sulfuric acid, the net ionic equation is:

$$Zn_{(s)} + 2H^+_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + H_{2(g)}$$



- In this reaction, zinc atoms form zinc cations. Each atom **loses two electrons**:

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

  - The electrons **balance the positive charge** of the metal ion.
  - The **lost** electrons are placed on the **right side** of the arrow.
  - Is this process oxidation or reduction?

- Now consider what is happening to hydrogen. Two hydrogen ions react to form one hydrogen molecule. In this process the ions **gain two electrons**.

$$2H^{+} + 2e^{-} \longrightarrow H_2$$

  - The electrons lost by zinc atoms are transferred to the hydrogen ions.
  - The **gained** electrons are shown on the **left side** of the arrow, and the charge of the reactants is the same as the charge of the products.
  - Is this process oxidation or reduction?

☐ Deriving half-equations from the full formula equation

- The process for determining the half-equations from the full formula equation is shown below:

1. Write a balanced full formula equation including states.

↓

2. Obtain the full ionic equation by splitting up aqueous ionic and acid species into charged ions.

↓

3. Obtain the net ionic equation by removing spectator ions.

↓

4. Separate the net ionic equation into two half-equations (one for each reactant) and add electrons to balance charges.

- Consider the following reaction:

$$2Na_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_{2(g)}$$

– Write the full ionic equation.

– Write the net ionic equation.

– What has happened to sodium in this reaction? Write the half-equation.

– What has happened to the hydrogen ions? Write the half-equation.

- Consider the following displacement reaction:

$$Ca_{(s)} + 2AgNO_{3(aq)} \longrightarrow Ca(NO_3)_{2(aq)} + 2Ag_{(s)}$$

  — Write the full ionic equation.

  — Write the net ionic equation.

  — Write the half-equations and label them as oxidation or reduction.

Combining half-equations

- Half-equations can also be combined to make the net ionic equation for a reaction.

- The process is shown below:

1. Write balanced half-equations including states. Ensure the number of atoms and ions are balanced and add electrons to balance the charges.

↓

2. Multiply the half-equations by suitable coefficients so that the electrons involved are the same (balance the electrons).

↓

3. Obtain the net ionic equation by adding the equations together and cancelling out electrons.

- Balance the pairs of half-equations below, classify them as oxidation or reduction, then combine them to form a balanced net ionic equation.

— $Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)}$

$Pb^{2+}_{(aq)} \longrightarrow Pb_{(s)}$

1

— $Ag^{+}_{(aq)} \longrightarrow Ag_{(s)}$

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

2

— $Ni_{(s)} \longrightarrow Ni^{2+}_{(aq)}$

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

3

— $Na_{(s)} \longrightarrow Na^{+}_{(aq)}$

$H_2O_{(l)} \longrightarrow OH^{-}_{(aq)} + H_{2(g)}$

4

— $FeO_4{}^{2-}{}_{(aq)} + H^{+}_{(aq)} \longrightarrow Fe^{3+}_{(aq)} + H_2O_{(l)}$

$F^{-}_{(aq)} \longrightarrow F_{2(g)}$

5

— CHALLENGE

$PbSO_{4(s)} + H_2O_{(l)} \longrightarrow PbO_{2(s)} + H^{+}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$

$ThO_{2(s)} + H^{+}_{(aq)} \longrightarrow Th_{(s)} + H_2O_{(l)}$

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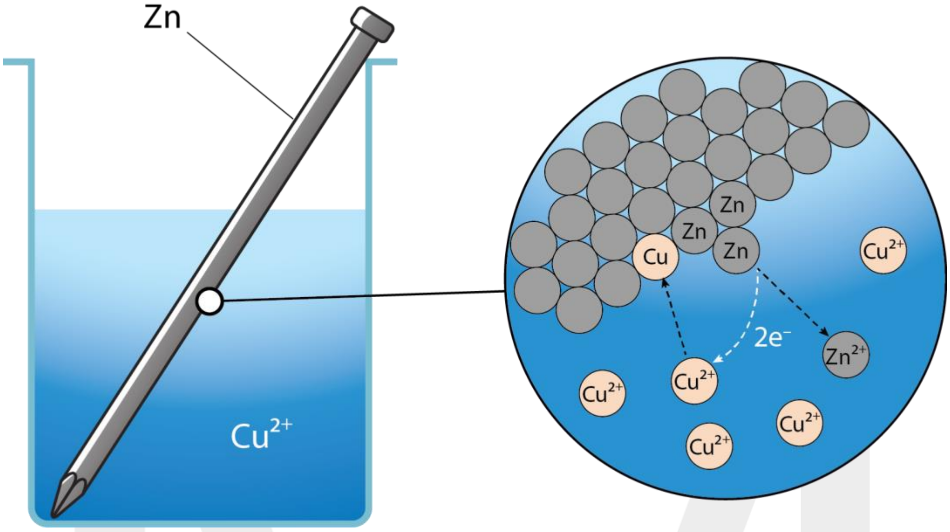
Concept Check 1.1

The diagram below shows the reaction between a zinc nail and a copper(II) solution.



(a) Which species is oxidised? Write the oxidation half-equation.

(b) Which species is reduced? Write the reduction half-equation.

(c) Add the oxidation and reduction half-equations to form the net ionic equation.

(d) What three macroscopic observations would you see taking place as the reaction proceeds?[7]

2. Oxidation states

☐ Oxidation numbers (oxidation states)

- Because redox reactions involve the gain and loss of electrons from atoms or ions, it is useful to determine which atoms specifically gain or lose electrons.
- Oxidation numbers (or oxidation states) indicate the **effective charge of an atom**, whether alone or in a compound.
- They are a useful concept for balancing equations and determining which species are oxidised and which are reduced.
- To determine oxidation number, follow these rules:

Rules	Examples
1. A pure element has an oxidation state of 0.	Na is $Na_{(s)}$, H in $H_{2(g)}$, Br in $Br_{2(l)}$
2. The oxidation state of an element in a monatomic ion equals the charge of that ion.	Oxidation state of Al in Al^{3+} is +3. Oxidation state of N in N^{3-} is -3.
3. The sum of all oxidation states of elements in a compound is 0.	In Al_2O_3, the oxidation state of O is -2, the oxidation state of Al is +3, overall: [3(-2)] + [2(+3)] = 0
4. The sum of all oxidation states of elements in a polyatomic ion is the charge of that ion.	In MnO_4^-, the oxidation state of O is -2, the oxidation state of Mn is +7, so that: (+7) + 4(-2) = -1
5. The oxidation state of hydrogen in its compounds is +1 (except for metal hydrides where it is -1).	Oxidation state of H in HCl or CH_4 is +1. Oxidation state of H in LiH is -1.
6. The oxidation state of oxygen in its compounds is -2 (except for peroxides where it is -1).	Oxidation state of O in H_2O is -2. Oxidation state of O in HNO_3 is -2.
7. In covalent compounds that don't have hydrogen or oxygen, the more electronegative element is assigned the negative oxidation state.	Oxidation state of Cl in PCl_3 is -1.

- For the following compounds and ions, determine the oxidation states of the underlined atom:

Species	Oxidation State	Species	Oxidation State
$\underline{Cu}$		$\underline{O}_2$	
$\underline{Fe}^{3+}$		$H\underline{F}$	
$\underline{C}O_2$		$\underline{Mn}O_4^-$	
$H\underline{N}O_3$		$\underline{Au}F_5$	
$\underline{V}_2O_5$		$\underline{Cr}_2O_7^{2-}$	
$\underline{Cl}F_5$		$\underline{Os}OF_5$	

- Oxidation states can be used to determine which species is oxidised (or reduced) since the oxidation state changes when species lose (or gain) electrons.
  - If an oxidation state **decreases** (becoming more negative), the species has been **reduced.**
  - If an oxidation number **increases** (becoming more positive), the species has been **oxidised.**

- For the reaction below, determine the change in the oxidation states to determine the species that had been oxidised and the species that has been reduced.

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

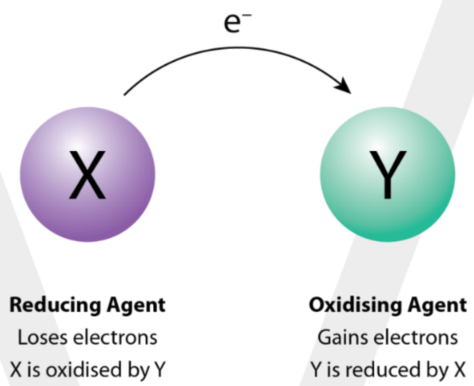
  - Species oxidised:

  - Species reduced:

☐ Oxidant and reductant

- An atom, molecule, or ion that causes another substance to be oxidised is called an **oxidising agent** or **oxidant**.
  - The oxidation state of the other substance increases.
  - The oxidising agent **is itself reduced**, and its oxidation state **decreases**.

- Similarly, an atom, molecule, or ion that causes another substance to be reduced is called a **reducing agent** or **reductant**.
  - The oxidation state of the other substance decreases.
  - The reducing agent **is itself oxidised**, and its oxidation state **increases**.



- Because oxidation and reduction always occur together, each redox reaction includes both an oxidising agent and a reducing agent.

- For each reaction below, write the half-equations and identify the oxidant and reductant.
  - $Mg_{(s)} + 2H^{+}_{(aq)} \longrightarrow Mg^{2+}_{(aq)} + H_{2(g)}$ 8

Oxidation: ______________________________

Reduction: ______________________________

Oxidant: ________________ Reductant: ________________

  - $Cl_{2(g)} + 2Br^{-}_{(aq)} \longrightarrow 2Cl^{-}_{(aq)} + Br_{2(l)}$ 9

Oxidation: ______________________________

Reduction: ______________________________

Oxidant: ________________ Reductant: ________________

- The activity series is a list of metals in the order of **decreasing strength** as reducing agents.
  - Metals rarely have a negative oxidation state, and therefore elemental metals are reducing agents in most of their reactions.
  - The **active metals** in Group 1, and some metals in Group 2 (Ca, Ba, Sr) are **very strong** reducing agents.

- The oxidised and reduced species that appear in a half-equation are sometimes called a **redox couple**.
  - In the above equation, the redox couples are $Zn^{2+}_{(aq)}$ / $Zn_{(s)}$ and $Cu^{2+}_{(aq)}$ / $Cu_{(s)}$.

Concept Check 2.1

Given the equation: $5Pb_{(s)} + 2MnO_4^-{}_{(aq)} + 16H^+{}_{(aq)} \longrightarrow 5Pb^{2+}{}_{(aq)} + 2Mn^{2+}{}_{(aq)} + 8H_2O_{(l)}$

(a) Identify the oxidation state of each element in each species.[10]

Element	Oxidation State
Pb in $Pb_{(s)}$	
Mn in $MnO_4^-{}_{(aq)}$	
O in $MnO_4^-{}_{(aq)}$	
H in $H^+{}_{(aq)}$	

Element	Oxidation State
Pb in $Pb^{2+}{}_{(aq)}$	
Mn in $Mn^{2+}{}_{(aq)}$	
H in $H_2O_{(l)}$	
O in $H_2O_{(l)}$	

(b) Identify the oxidising and reducing agents.[11]

Concept Check 2.2

For the following redox reactions, write the net ionic equation and the two half-equations, and label the half-equations as oxidation or reduction.

(a) Tin(II) ions reduce iron(III) to iron(II). Each tin(II) ion loses two electrons in this process.

(b) Aqueous bromine reacts with iodide ions to form bromide ions and iodine solution.[12]