

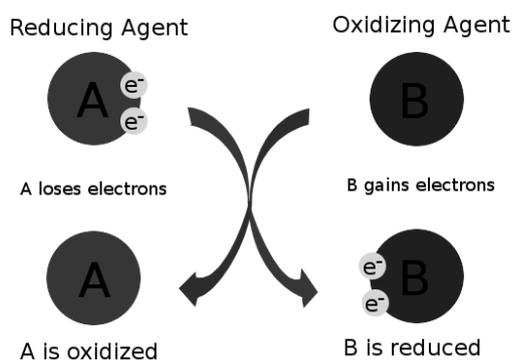
WASHINGTON LATIN PUBLIC CHARTER SCHOOL

CHEMISTRY 2019-20

UNIT 5B

CHEMICAL REACTIONS II – OXIDATION AND REDUCTION

Help with distance learning lessons



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Key words: oxidation, reduction, half-equation, redox reaction, oxidizing agent, reducing agent, alkali metal, displacement reaction, reactivity series, galvanic cell, voltaic cell, anode, cathode, battery, electrolyte, electrolytic cell

Lesson 1 – Introduction to Oxidation and Reduction

- Oxidation and reduction are best defined in terms of electron transfer:

Oxidation is the loss of electrons

When a species loses electrons it is said to be **oxidised**

Eg A sodium atom (Na) can lose one electron to become a sodium ion (Na⁺)
This process can be represented as follows: $\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$

Eg A magnesium atom (Mg) can lose two electrons to become a magnesium ion (Mg²⁺)
This process can be represented as follows: $\text{Mg} \rightarrow \text{Mg}^{2+} + \text{e}^-$

Eg An iodide ion (I⁻) can lose an electron to become an iodine atom (I)
This process can be represented as follows: $\text{I}^- \rightarrow \text{I} + \text{e}^-$

Reduction is the gain of electrons

When a species gains electrons it is said to be reduced

Eg A Cl atom can gain one electron to form a Cl⁻ ion
This process can be represented as follows: $\text{Cl} + \text{e}^- \rightarrow \text{Cl}^-$

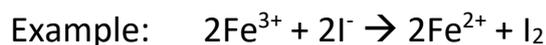
Eg An O atom can gain two electrons to become an oxide ion (O²⁻)
This process can be represented as follows: $\text{O} + 2\text{e}^- \rightarrow \text{O}^{2-}$

Eg An Fe³⁺ ion can lose one electron to become a Fe²⁺ ion
This process can be represented as follows: $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$

- Remember: LEO GER (loss of electrons is oxidation, gain of electrons is reduction) or OIL RIG (oxidation is loss, reduction is gain)
- Equations such as those shown above, which show the gain or loss of electrons by a species, are known as **half-reactions** or **half-equations**
- **A redox reaction is a reaction involving the transfer of electrons**
During redox reactions, one species loses electrons and gets oxidised, whilst another gains electrons and gets reduced.

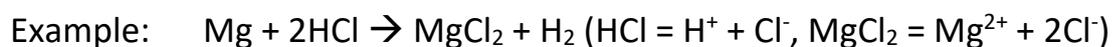
Example: $\text{Na} + \text{Cl} \rightarrow \text{NaCl}$ (NaCl contains Na⁺ and Cl⁻)
During this reaction, Na becomes Na⁺ - it loses an electron so is oxidized
During this reaction, Cl becomes Cl⁻ - it gains an electron so is reduced

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During this reaction, 2I^- becomes I_2 - they lose electrons so are oxidized

During this reaction, Fe^{3+} becomes Fe^{2+} - it gains electrons so is reduced



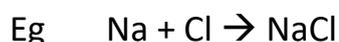
During this reaction, Mg becomes Mg^{2+} - it loses an electron so is oxidized

During this reaction, 2H^+ becomes H_2 – they gain electrons so are reduced

The Cl^- stays as Cl^- – it is neither oxidized nor reduced

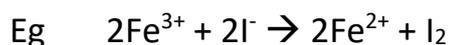
Lesson 2 – Oxidizing and Reducing Agents

- In a redox reaction, the atom which is reduced is accepting electrons from the other species and causing it to be oxidised; the atom being reduced is therefore an **oxidising agent**; an oxidising agent is an **electron acceptor**; it causes the atom it reacts with to be oxidized
- The atom which is oxidized is donating electrons to another atom and causing it to be reduced; the atom being oxidized is therefore a **reducing agent**; a reducing agent is an electron donor; it causes the atom it reacts with to be reduced
- A redox reaction can thus be described as a transfer of electrons from a reducing agent to an oxidising agent



In this reaction Na is oxidized ($\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$) so Na is the reducing agent

In this reaction Cl is reduced ($\text{Cl} + \text{e}^- \rightarrow \text{Cl}^-$) so Cl is the oxidizing agent



In this reaction I^- is oxidized ($2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$) so I^- is the reducing agent

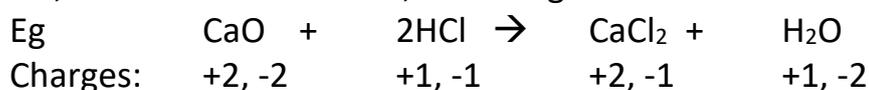
In this reaction Fe^{3+} is reduced ($\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$) so Fe^{3+} is the oxidizing agent



In this reaction Mg is oxidized ($\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$) so Mg is the reducing agent

In this reaction H^+ is reduced ($2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$) so H^+ is the oxidizing agent

- Note - not all reactions are redox reactions; in acid-base reactions, it is H^+ ions, not electrons, which are transferred; the charges on the atoms do not change:



If the charges on the atoms do not change, the reaction is NOT a redox reaction

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- Oxidizing agents are useful as disinfectants and in treating wounds, as the oxidation of microbes tends to kill them:
 - Swimming pools contain Cl_2 , chlorox contains NaClO and I_2 is used to sterilize wounds
 - Cl_2 , NaClO and I_2 are all oxidizing agents
- Reducing agents are often added to food to stop it going bad (due to oxidation); they are often referred to as “anti-oxidants”
 - Na_2SO_3 is a reducing agent; it is often added to wine (to prevent the alcohol from being oxidized into vinegar) and to some food (so they don't go bad as quickly)

Lesson 3 – Reaction of Metals with Acid and Water

- Most metals react with acids to make a **salt** and **hydrogen**
Eg $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$ (salt = magnesium chloride)
 $\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2$ (salt = zinc sulfate)
- This reaction is an example of a **redox** reaction:
 - the metal is oxidised from 0 to +2 (in the above reactions):
 $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$ or $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$
 - the H^+ ion in the acid is reduced from +1 to 0 (in hydrogen): $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$
 - the metal is the reducing agent and the H^+ is the oxidizing agent
 - the chloride and sulfate ions are neither oxidized or reduced; they remain in solution, unchanged, and can be omitted from the equation (although you can keep them in if you want)
- The rate of reaction between metals and acids varies considerably, and many metals do not react with acids; it depends on the reactivity of the metal, which depends on the ability of the metal to give electrons to the H^+ ions; some metals lose their electrons much more easily than others
- Water also contains H^+ , ($\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$) although the H^+ ions are present in smaller concentrations than acids; some very reactive metals can react with the H^+ in water:
Eg $2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2$
- In the above reaction, the Na is oxidised from 0 to +1 and the H is reduced from +1 to 0; the products are a metal hydroxide and hydrogen; if the hydroxide is soluble in water, the resulting solution will contain OH^- ions and will therefore be alkaline

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- The only metals which are reactive enough to do this are the Group I metals and some Group II metals; these metals therefore react with water to produce alkaline solutions, and for this reason Group I metals are often called “**the alkali metals**”

Lesson 4 – Displacement Reactions and the Reactivity Series

- A metal can displace another metal from any compound, provided that the other metal is less reactive (and therefore has a lower tendency to be oxidized); these reactions are known as **displacement reactions**

Here is an example of a displacement reaction: $\text{Zn} + \text{CuO} \rightarrow \text{ZnO} + \text{Cu}$

Charges: 0 +2, -2 +2, -2 0

- the more reactive metal is zinc; it is oxidised from 0 to +2 ($\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$)
 - the less reactive metal is Cu; is reduced from +2 to 0 ($\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$)
 - the oxide ion is not directly involved in the reaction (its charge does not change)
- The above reaction only happens because zinc is more reactive than copper; the reverse reaction cannot happen, as copper is less reactive than zinc and so cannot displace zinc from its compounds: $\text{Cu} + \text{ZnO} \rightarrow$ (no reaction)
 - Reactions between metals and acids can also be described as displacement reactions; the metal is only able to displace hydrogen from acids if the metal is more reactive than hydrogen:
Eg $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$
zinc is more reactive than hydrogen so can displace H from acids
 $\text{Cu} + 2\text{HCl} \rightarrow$ no reaction
copper is less reactive than hydrogen so cannot displace H from acids

- Many metals are found in nature (ie in their ores) as metallic oxides; during their extraction from their ores, they have to be converted from the metal oxide to the metal, this is often carried out by reacting them with carbon or hydrogen in a displacement reaction:

Eg zinc can be extracted from its ore by reaction with carbon:



This reaction happens because carbon is more reactive than zinc; copper, iron, tin and lead can also be extracted in this way

Aluminium cannot be extracted in this way, because carbon is less reactive than aluminium: $\text{Al}_2\text{O}_3 + \text{C} \rightarrow$ no reaction

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- By analysing the results of metal displacement reactions, including the reactions of metals with acids, metals with water, and metals, carbon and hydrogen with metal compounds, it is possible to rank all metals (and some electropositive non-metals such as carbon and hydrogen) in order of reactivity; this is called the **reactivity series**
- The more reactive the metal, the greater its tendency to be oxidized, the more effective it is as a reducing agent, and the more difficult it is to extract from its compounds

Metal	Notes
potassium	very reactive – stored under oil – reacts violently with water
sodium	very reactive – stored under oil – reacts violently with water
calcium	rapidly with water
magnesium	reacts rapidly with acids, very slowly with water
aluminium (carbon)	very slow reactions with acids, water and air at room temperature*
zinc	reacts steadily with acids
iron	rusts slowly in air and water
tin	reacts very slowly with acids
lead	reacts very slowly with acids
(hydrogen)	
copper	no reaction with acids; very slow reaction with air and water
silver	no reactions with acids, air or water
gold	no reactions with acids, air or water

- *Note the low apparent reactivity of aluminium, given its high position in the reactivity series; aluminium is a highly reactive metal, but quickly forms a layer of its oxide (Al_2O_3) on its surface; this oxide is very stable and unreactive, and sticks to the surface of the aluminium, significantly slowing down its reactions; for this reason, aluminium metal is relatively resistant to oxidation and is widely used in everyday materials

Lesson 5 – Using redox reactions to generate electricity

(a) Introduction to Galvanic Cells

- A galvanic (or voltaic) cell is a device which uses a redox reaction to generate electricity
- All galvanic cells contain two compartments, called half-cells;
 - in one compartment, atoms get oxidised (so they lose electrons)
 - in the other compartment, atoms get reduced (so they gain electrons)
 - the electrons flow from one compartment to the other through an electrical wire, creating electricity
 - both compartments contain a conducting rod dipped into a solution containing ions; these rods are called **electrodes** and they are both connected to the wire
 - the electrode at which reduction takes place is called the **cathode**
 - the electrode at which oxidation takes place is called the **anode**
 - the electrodes act as the terminals of the cells; the cathode is the positive terminal and the anode is the negative terminal
- The simplest example of a galvanic cell is a Daniell cell, which uses the reaction between zinc and copper sulfate to generate electricity
- Consider the reaction:

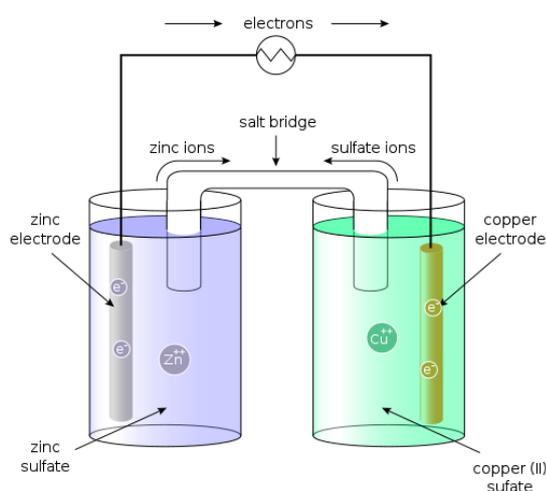
Zn	+	CuSO ₄ (aq)	→	ZnSO ₄ (aq)	+	Cu(s)
charges:		0		+2		+2
			→			0

In this reaction:

- the Zn atoms is oxidised to Zn²⁺: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$
- therefore the zinc is the anode of the cell (it is negative)
- after being oxidized, the Zn atoms leave the electrode and go into the solution as Zn²⁺ ions
- therefore the Zn electrode gets gradually smaller
- the Cu²⁺ ions are reduced to Cu: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$
- therefore the copper is the cathode of the cell (it is positive)
- after being reduced, the Cu²⁺ ions leave the solution and attach themselves to the electrode as Cu atoms
- therefore the Cu electrode gets gradually bigger
- the net ionic equation is: $\text{Zn} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu}(\text{s})$

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- this is a picture of the cell:



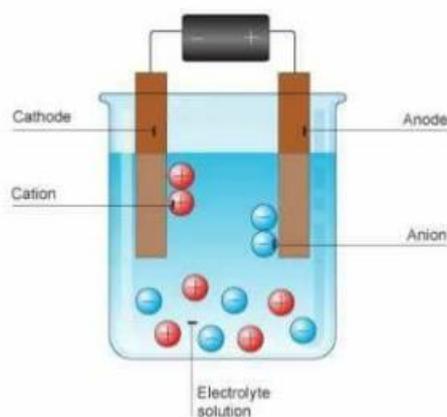
(b) Examples of useful batteries

- Batteries are devices containing one more more cells linked together; the greater the voltage desired, the greater the number of cells which need to be linked up in a battery
- Some batteries are re-chargeable; when this happens, the reverse reaction takes place and the original reactants are restored
- There are many important examples of Galvanic cells; they have a variety of uses and represent a significant area of current scientific research; three common examples are highlighted below
- The **lead-acid battery** is used in cars and has several other applications, it is rechargeable and is able to withstand a high current flow for a short time; however it is very heavy
 - the reaction taking place is $\text{Pb} + \text{PbO}_2 + 2\text{H}_2\text{SO}_4 \rightarrow 2\text{PbSO}_4 + 2\text{H}_2\text{O}$
- The **alkaline battery** is the most commonly used household battery; it is cheap and portable, although it is non-rechargeable
 - the reaction taking place is $2\text{MnO}_2 + \text{Zn} \rightarrow \text{Mn}_2\text{O}_3 + \text{ZnO}$
- The **lithium-ion battery** is widely used in cellphones and other electronic devices; it is light, rechargeable and powerful, but the chemicals needed to make the batteries are very expensive
 - the reaction taking place is $\text{CoO}_2 + \text{Li} \rightarrow \text{LiCoO}_2$

Lesson 6 – Electrolysis

(a) Principles of Electrolysis

- If two electrodes are dipped into a liquid containing ions (called an electrolyte) connected to a power supply, the anions in the solution will move towards the positively charged electrode (known as the anode) and the cations in the solution will move towards the negatively charged electrode (known as the cathode)
- Once the anions reach the anode, they are oxidised, usually to a non-metallic element (eg Cl^- ions are oxidised to Cl_2)
- Once the cations reach the cathode, they are reduced, usually to a metallic element or hydrogen (Cu^{2+} ions are reduced to Cu and H^+ ions are reduced to H_2)
- The use of electricity to separate ionic compounds into elements is known as **electrolysis**
- The apparatus in which electrolysis takes place is called an **electrolytic cell**; electrolytic cells are used to convert electrical energy into chemical energy; this is the opposite of a Galvanic cell, in which chemical energy is converted into electrical energy; an example of an electrolytic cell is shown below:



(b) Electrolysis of molten ionic compounds

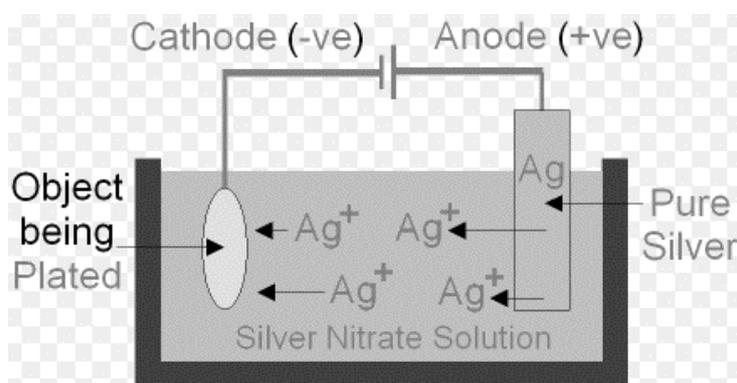
- Ionic compounds conduct electricity when molten as the ions are free to move
 - for very reactive metals (such as sodium and potassium) and very reactive non-metals (such as fluorine) this is often the only way to make the pure element, as they are too reactive to be displaced chemically
 - **eg molten sodium chloride** contains Na^+ and Cl^- ions
 - During electrolysis, the Na^+ ions move to the cathode where they are reduced, and the Cl^- ions move to the anode where they are oxidised
 - **sodium** is produced at the cathode and **chlorine** is produced at the anode

(c) Electrolysis of aqueous ionic compounds

- Ionic compounds also conduct electricity when dissolved in water, as the ions are able to move; it is therefore possible to electrolyse aqueous solutions; this is much cheaper (and safer) than electrolysing molten ionic compounds as the ionic compound does not need to be heated
- Aqueous solutions, however, contain H^+ ions and OH^- ions due to the water which is also present, as well as any ions which dissolve in water; this means that there is often a choice of cation to be reduced and a choice of anion to be oxidised
- If the solution contains the cation of a metal which is more reactive than hydrogen (eg Na^+ , Fe^+ or Zn^{2+}), then **hydrogen (H_2)** is produced at the cathode and the metal ion will remain in solution
- If the solution contains the cation of a metal which is less reactive than hydrogen (eg Cu^{2+} or Ag^+), then **the metal** is produced at the cathode and the hydrogen ions remain in solution
- Most anions are more difficult to oxidise than OH^- ions, so **oxygen (O_2)** is usually produced at the anode
- Br^- and I^- are easier to oxidise than OH^- ions, so in aqueous solutions containing these ions, the halogen (**bromine (Br_2)** or **iodine (I_2)**) is produced at the anode, and not oxygen
- Cl^- and OH^- ions have a similar tendency to be oxidised, so if an aqueous solution containing Cl^- ions is electrolysed, the gas produced at the anode could either be **oxygen (O_2)** or **chlorine (Cl_2)**, depending on the concentration of the solutions

(d) Electroplating

- **Electroplating** is the use of electrolysis to cover the surface of one metal with a thin layer of another metal; it can be used either to protect or to decorate the metal
 - for example, to electroplate a copper vessel with a thin layer of silver, the vessel should be used as the cathode in the electrolytic cell and the electrolyte should contain silver ions
 - the Ag^+ ions move to the cathode and are reduced to Ag, covering the vessel with a thin layer of silver
 - electroplating with silver is commonly used to improve the appearance of cutlery, crockery and jewellery



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