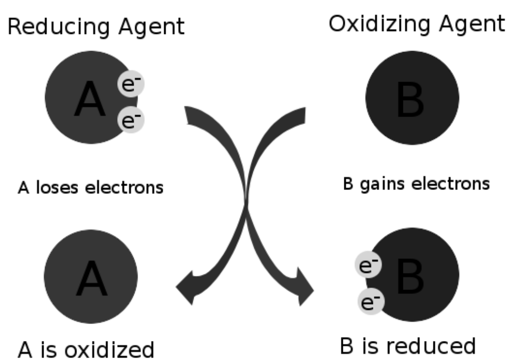
WASHINGTON LATIN PUBLIC CHARTER SCHOOL

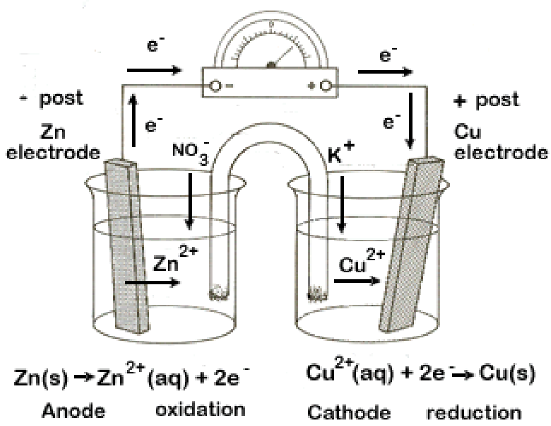
HONORS CHEMISTRY 2019-20

UNIT 5B

CHEMICAL REACTIONS II – OXIDATION AND REDUCTION

(Lessons 1 – 3)





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Key words: oxidation, reduction, oxidation number, oxidising agent, reducing agent, half-equation, redox reaction

Lesson 1 – What is oxidation, what is reduction and what are half-equations?

1) Oxidation and Reduction

a) Definitions of 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: $Na \rightarrow Na^+ + e^-$
    Eg A magnesium atom (Mg) can lose two electrons to become a magnesium ion (Mg^{2+})
    This process can be represented as follows: $Mg \rightarrow Mg^{2+} + e^-$
    Eg Two iodide ions (I^-) can each lose an electron to become an iodine molecule (I_2)
    This process can be represented as follows: $2I^- \rightarrow I_2 + 2e^-$

  - **Reduction is the gain of electrons**; when a species gains electrons it is said to be reduced
    Eg Each Cl atom in Cl_2 can gain one electron to form two Cl^- ions
    This process can be represented as follows: $Cl_2 + 2e^- \rightarrow 2Cl^-$
    Eg Each O atom in O_2 can gain two electrons to become an oxide ion (O^{2-})
    This process can be represented as follows: $O_2 + 4e^- \rightarrow 2O^{2-}$
    Eg An Fe^{3+} ion can lose one electron to become a Fe^{2+} ion
    This process can be represented as follows: $Fe^{3+} + e^- \rightarrow Fe^{2+}$

- 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**

b) Naming ionic compounds

- Most ionic compounds are named by stating the cation followed by the anion (eg NaCl - sodium chloride)

- Atoms in groups 1 only form +1 ions and atoms in group 2 only form +2 ions, so it is not necessary to specify the charge on the atom when naming these compounds

- Some atoms, especially those in the d-block, form more than one stable cation and can therefore form different compounds with the same anion:
  Eg Fe can form Fe^{2+} or Fe^{3+} ions, so can form $FeCl_2$ or $FeCl_3$
  Eg Cu can form Cu^+ or Cu^{2+} ions, so can form Cu_2O or CuO

  To distinguish between these different compounds by name, the charge on the cation must be specified; this is done by stating the charge in Roman numerals in parenthesis after the name of the cation:
  Fe^{2+} is iron (II) so $FeCl_2$ is iron (II) chloride
  Fe^{3+} is iron (III) so $FeCl_3$ is iron (III) chloride
  Cu^+ is copper (I) so Cu_2O is copper (I) oxide
  Cu^{2+} is copper (II) so CuO is copper (II) oxide
  Other examples: PbO_2 is lead (IV) oxide (because it contains Pb^{4+} ions)
  V_2O_5 is vanadium (V) oxide (because it contains V^{5+} ions)

  Note: you do NOT need to state the charge when naming compounds of atoms in Group 1, Group 2 or aluminum, as these atoms only form one stable ion (MgO is magnesium oxide)

Lesson 2 – What are oxidation numbers?

c) Oxidation numbers

- The concepts of oxidation and reduction can be applied to non-ionic compounds using the concept of **oxidation numbers**

- **The oxidation number of an atom is the charge that would exist on an individual atom if the bonding were completely ionic**

- In simple ions, the oxidation number of the atom is the same as the charge on the ion:
  - Na^+, K^+, H^+ all have an oxidation number of +1
  - Mg^{2+}, Ca^{2+}, Pb^{2+} all have an oxidation number of +2
  - Cl^-, Br^-, I^- all have an oxidation number of -1
  - O^{2-}, S^{2-} all have an oxidation number of -2

- In molecules and compounds, the sum of the oxidation numbers on the atoms is zero
  - In SO_3; oxidation number of S = +6, oxidation number of each O = -2 (+6 + 3(-2) = 0)
  - H_2O_2; oxidation number of each H = +1, oxidation number of each O = -1 (2(+1) + 2(-1) = 0)
  - SCl_2; oxidation number of S = +2, oxidation number of each Cl = -1 (+2 + 2(-1) = 0)

- In elements, the oxidation number of each atom is zero
  - In Cl_2, S, Na and O_2 all atoms have an oxidation number of zero

- In polyatomic ions, the sum of the oxidation numbers on the atoms is equal to the overall charge on the ion.
  - In SO_4^{2-}; oxidation number of S = +6, oxidation number of O = -2 +6 + 4(-2) = -2
  - In ClO^-; oxidation number of Cl = +1, oxidation number of O = -2 +1 + (-2) = -1
  - In NH_4^+; oxidation number of N = -3, oxidation number of H = +1 -3 +4(1) = +1

UNIT 5B – OXIDATION AND REDUCTION

- Many atoms, such as S, N and Cl, can exist in a variety of oxidation states and so it is not possible to assume that these atoms have a particular oxidation number in their compounds and ions; oxidation numbers can be predicted by following a number of basic rules:
  - All group I atoms always adopt the +1 oxidation state in their compounds (eg NaCl)
  - All group II atoms adopt the +2 oxidation state in their compounds (eg MgO)
  - Aluminum always adopts the +3 oxidation state in its compounds (eg Al_2O_3)
  - Fluorine always adopts the -1 oxidation state in its compounds (eg HF)
  - Hydrogen usually adopts the +1 oxidation state in its compounds and ions (eg HCl), but when it is bonded to a metal, silicon or boron it adopts the -1 oxidation state; these compounds are called **hydrides** (eg BH_3 – boron trihydride)
  - Oxygen usually adopts the -2 oxidation state in its compounds (eg H_2O, SO_4^{2-}); in some cases, when bonded to a group I or group II metal or hydrogen, it can adopts the -1 oxidation state; these compounds are called **peroxides** (eg H_2O_2 – hydrogen peroxide); with fluorine, oxygen adopts an oxidation number of +2
  - The oxidation numbers of all other atoms in their compounds can vary

Example 1: What are the oxidation numbers of C and O in carbon monoxide, CO?
Solution: The oxidation number of O is -2, as C is not one of the exceptions to the oxidation number of O; therefore the oxidation number of C must be +2, because the overall charge on a molecule is zero **So ON of C = +2, ON of O = -2**

Example 2: What are the oxidation numbers of N and O in the nitrate ion, NO_3^-?
Solution: The oxidation number of O is -2, as N is not one of the exceptions to the oxidation number of O; therefore the oxidation number of N must be +5, because the overall charge on the ion = -1 **So ON of N = +5, ON of O = -2**

- During oxidation and reduction, the oxidation numbers of atoms change:

  - If an atom is oxidized, its oxidation number increases (ie it becomes more +ve or less –ve)
  - If an atom is reduced, its oxidation number decreases (ie it becomes less +ve or more –ve)

Oxidation	Loss of electrons	Increase in oxidation number
Reduction	Gain of electrons	Decrease in oxidation number

Example: During a reaction, PbO_2 is converted into PbO. What has happened to the Pb in this reaction?
Answer: the ON of Pb in PbO_2 = +4; the ON of Pb in PbO = +2; so the Pb has been reduced

Example: During a reaction, CH_4 is converted into CO_2. What has happened to the C in this reaction?
Answer: the ON of C in CH_4 = -4; the ON of C in CO_2 = +4; so the C has been oxidised

- Oxidation and reduction processes involving polyatomic ions or molecules can also be expressed as half-equations, but they are more complex and often involve H^+ and H_2O:
  Eg The reduction of PbO_2 to Pb^{2+} can be written $PbO_2 + 4H^+ + 2e^- \rightarrow Pb^{2+} + 2H_2O$
  Eg The oxidation of SO_3^{2-} to SO_4^{2-} can be written $SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2e^-$

- You will not be asked to derive these half-equations, but you will be expected to use them and to recognise them as oxidation or reduction

Lesson 3 – What are redox reactions, and what are oxidising and reducing agents?

d) Redox reactions

- Half-equations consider gain and loss of electrons, but in fact electrons cannot be created or destroyed; they can only be transferred from species to species; the gain of electrons by one species necessarily involves loss of electrons by another; oxidation and reduction thus always occur simultaneously; an oxidation is always accompanied by a reduction and vice versa; any reaction consisting of the oxidation of one species and the reduction of another is known as a **redox** reaction

- A redox reaction can be described as a reaction which involves the **transfer of electrons from one species to another**

(i) Deriving equations for redox reactions

- A redox reaction can be derived by combining an oxidation half-equation with a reduction half-equation in such a way that the total number of electrons gained is equal to the total number of electrons lost; in some cases, H_2O and H^+ may appear on both sides of the reaction, which must then be simplified

  Eg $H_2SO_4 + 8H^+ + 8e^- \rightarrow H_2S + 4H_2O$ (reduction) and $2I^- \rightarrow I_2 + 2e^-$ (oxidation)
  - the oxidation half-equation must be multiplied by 4 to equate the electrons: $8I^- \rightarrow 4I_2 + 8e^-$
  - the two half-equations can then be added together:
  - $\mathbf{H_2SO_4 + 8H^+ + 8I^- \rightarrow H_2S + 4H_2O + 4I_2}$

  Eg $PbO_2 + 4H^+ + 2e^- \rightarrow Pb^{2+} + 2H_2O$ (reduction) and $SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2e^-$ (oxidation)
  - the number of electrons is the same in both half-equations so multiplying one of them is not required
  - adding the two half-equations together gives $PbO_2 + 4H^+ + SO_3^{2-} + H_2O \rightarrow Pb^{2+} + 2H_2O + SO_4^{2-} + 2H^+$
  - removing $2H^+$ and H_2O from both sides gives:
  - $\mathbf{PbO_2 + 2H^+ + SO_3^{2-} \rightarrow Pb^{2+} + H_2O + SO_4^{2-}}$

- A redox reaction can be derived by combining an oxidation half-equation with a reduction half-equation in such a way that the total number of electrons gained is equal to the total number of electrons lost; in some cases, H_2O and H^+ may appear on both sides of the reaction, which must then be simplified

(ii) Oxidising agents and reducing agents

- The species which is reduced is accepting electrons from the other species and thus causing it to be oxidised; the reduced species is therefore an **oxidising agent**; an oxidising agent is an **electron acceptor**; it causes the oxidation number of the other species to increase
  - H_2SO_4, Al^{3+} and Cl_2 are all oxidising agents

- The species which is oxidised is donating electrons to another species and thus causing it to be reduced; it is therefore a **reducing agent**; a reducing agent is an electron donor; it causes the oxidation number of the other species to decrease

- Na, O^{2-}, I^- and $S_2O_3^{2-}$ are all reducing agents

- A redox reaction can thus be described as a transfer of electrons from a reducing agent to an oxidising agent

- It is possible to identify what is oxidized and what is reduced (and hence the oxidizing agent and the reducing agent) in a chemical reaction by considering the changes in oxidation number of the different atoms:

Eg $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$
The ON of I decreases from 0 to -1 (I_2 is reduced)
the ON of S increases from +2 to +2.5 ($S_2O_3^{2-}$ is oxidized)
I_2 is the oxidizing agent; $S_2O_3^{2-}$ is the reducing agent

Eg $H_2SO_4 + 8HI \rightarrow H_2S + 4H_2O + 4I_2$
The ON of S decreases from +6 to -2 (H_2SO_4 is reduced)
the ON of I increases from -1 to 0 (HI is oxidized)
H_2SO_4 is the oxidizing agent, HI is the reducing agent

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

Eg	CaO	+	2HCl	$\rightarrow$	$CaCl_2$	+	H_2O
Charges:	+2, -2		+1, -1		+2, -1		+1, -2

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

- 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)

(iii) Disproportionation reactions

- In some reactions, the same species is simultaneously both oxidized and reduced:

Eg $2H_2O_2 \rightarrow 2H_2O + O_2$
The oxidation number of O increases from -1 (in H_2O_2) to 0 (in O_2)
The oxidation number of 0 decreases from -1 (in H_2O_2) to -2 (in H_2O)
So H_2O_2 is oxidizing and reducing itself
H_2O_2 is the oxidizing agent and the reducing agent

Eg $Na_2S_2O_3 + 2HCl \rightarrow S + SO_2 + 2H_2O$
The oxidation number of S increases from +2 (in $Na_2S_2O_3$) to +4 (in SO_2)
The oxidation number of S decreases from +2 (in $Na_2S_2O_3$) to 0 (in S)
So $Na_2S_2O_3$ is oxidizing and reducing itself
$Na_2S_2O_3$ is the oxidizing agent and the reducing agent

The simultaneous oxidation and reduction of the same species in a chemical equation is called **disproportionation**