Department of Mathematics and Natural Sciences
CHE 101: Introduction to Chemistry

Lecture 11

Content: Redox Reaction

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Oxidation-Reduction (Redox) Reactions

- Oxidation-Reduction (Redox) reactions are electron transfer reactions which are considered to be a part of everyday life.
- Most metallic and non-metallic elements are obtained from their ores(naturally occurring solid material) by either oxidation or reduction reactions. Many (but not all) important redox reactions take place in aqueous systems.
- ***Oxidation reaction:*** Half-reaction that refers to loss of electron/electrons

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

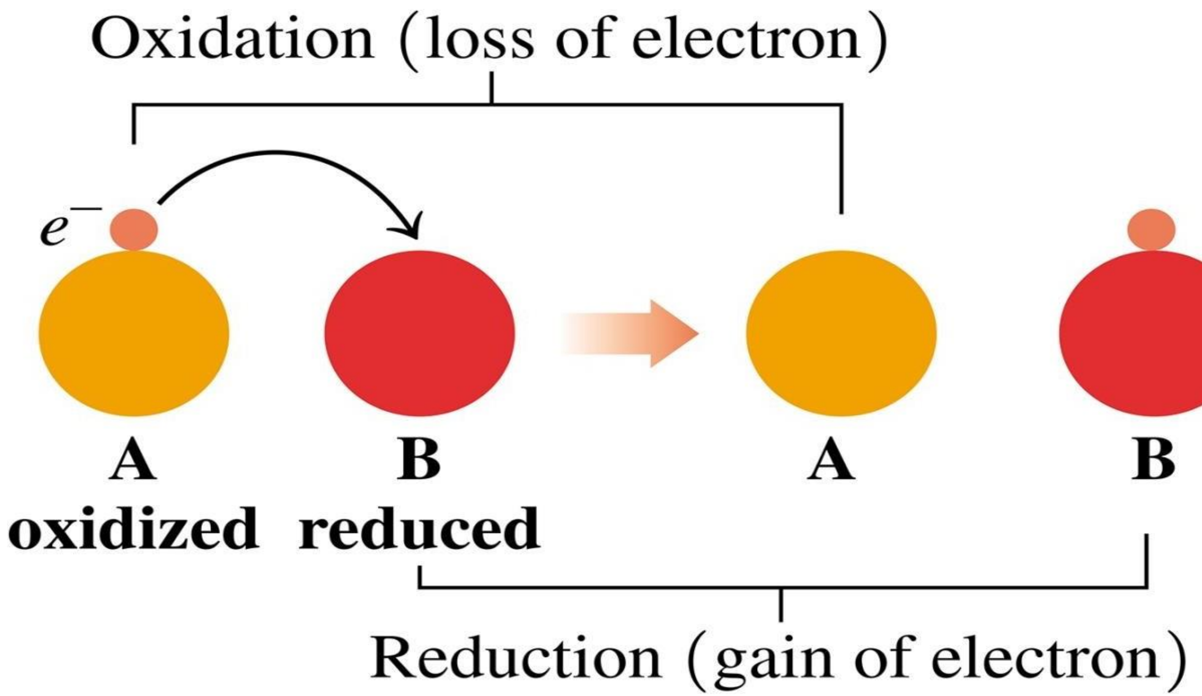
- ***Reduction reaction:*** Half-reaction that refers to gain of electron/electrons

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

- In the formation of CaO, Ca is oxidized, and itself acts as a reducing agent by giving up electrons, while oxygen is reduced, and itself acts as an oxidizing agent by accepting electrons.

Oxidation–Reduction

In an **oxidation–reduction reaction**, electrons are transferred from one substance to another.



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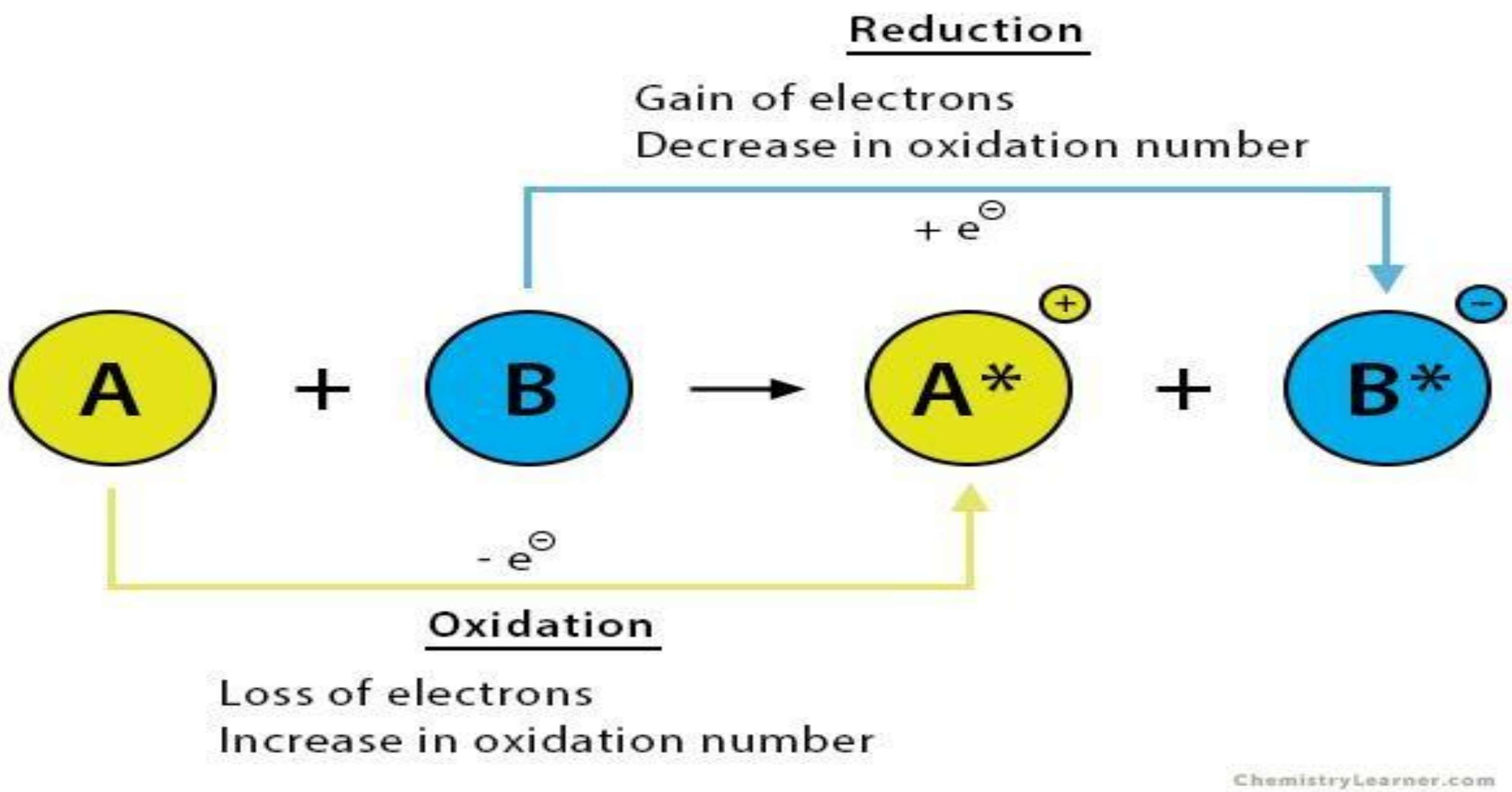
Oxidation Is Loss of electrons. Reduction Is Gain of electrons.

Redox Reaction



Oxidation-Reduction (Redox) Reactions

- Def'n: Reactions in which one or more electrons is shifted from one element to another
(Note: In all reactions discussed previously, atoms changed partners, but each atom kept all its electrons)

- **Oxidation** = loss of e^{-}, hence higher charge;
Reduction = gain of e^{-}, hence lower charge;
Redox = transfer of e^{-} (simultaneous oxidation and reduction)

Oxidation–Reduction

The green patina on the Statue of Liberty is due to the oxidation of copper metal as it forms a green solid, CuO.

$2Cu(s) \rightarrow 2Cu^{2+}(s) + 4e^{-}$ oxidation

$O_2(g) + 4e^{-} \rightarrow 2O^{2-}(s)$ reduction

$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$

Core Chemistry Skill Identifying Oxidized and Reduced Substances

Study Check

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Identify each of the following as oxidation or reduction:

A. $Sn(s) \rightarrow Sn^{4+}(aq) + 4e^{-}$

B. $Fe^{3+}(aq) + 1e^{-} \rightarrow Fe^{2+}(aq)$

C. $Cl_2(g) + 2e^{-} \rightarrow 2Cl^{-}(aq)$

Solution

Identify each of the following as oxidation or reduction:

A. $Sn(s) \rightarrow Sn^{4+}(aq) + 4e^{-}$ Oxidation

B. $Fe^{3+}(aq) + 1e^{-} \rightarrow Fe^{2+}(aq)$ Reduction

C. $Cl_2(g) + 2e^{-} \rightarrow 2Cl^{-}(aq)$ Reduction

Study Check

In light-sensitive sunglasses, UV light initiates an oxidation–reduction reaction.

$$2Ag^{+} + 2Cl^{-} \xrightarrow{\text{UV light}} 2Ag + Cl_2$$

A. Which reactant is oxidized?

B. Which reactant is reduced?

Solution

In light-sensitive sunglasses, UV light initiates an oxidation–reduction reaction.

$$2Ag^{+} + 2Cl^{-} \xrightarrow{\text{UV light}} 2Ag + Cl_2$$

A. Which reactant is oxidized? chloride ion, Cl^{-}

$$2Cl^{-} \rightarrow Cl_2 + 2e^{-}$$

B. Which reactant is reduced? silver ion, Ag^{+}

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

Characteristics of Oxidation and Reduction

TABLE 7.4 Characteristics of Oxidation and Reduction

Always Involves	May Involve
Oxidation	
Loss of electrons	Addition of oxygen
	Loss of hydrogen
Reduction	
Gain of electrons	Loss of oxygen
	Gain of hydrogen

Redox in Photosynthesis

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

Redox in Photosynthesis

Reduction

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

Oxidation

TABLE 7.4 Characteristics of Oxidation and Reduction

Always Involves	May Involve
Oxidation	
Loss of electrons	Addition of oxygen
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Gain of electrons	Loss of oxygen
	Gain of hydrogen

Oxidation Number (ON)

- It is the number charge assigned on an element or a species (elements, ions or molecules) during its loss or gain of electron.
- For the loss of an electron = a positive Oxidation Number (ON) is obtained
  $Na\ (g) \rightarrow Na^{+}\ (g) + e^{-}$
- For the gain of an electron = a negative ON is obtained

$Cl\ (g) + e^{-} \rightarrow Cl^{-}(g)$

- The numbers above the element symbols are the oxidation numbers. As can be seen in both the reactions the charges on the atoms on the reactant molecules are zero. On the other hand it is assumed that complete transfer (*loss or gain*) of electron have taken place on the atoms in the product molecules. The oxidation number (ON) reflects the number of electrons transferred.

Oxidation numbers are assigned to elements using these rules:

Rule 1: The oxidation number of an element in its free (uncombined) state is zero — for example, Al(s) or Zn(s). This is also true for elements found in nature as diatomic (two- atom) elements H2, O2, N2, Cl2, Br2

Rule 2: The oxidation number of a monatomic (one-atom) ion is the same as the charge on the ion, for example:

$$Na^{+1} = +1, S^{2-} = -2$$

Rule 3: The sum of all oxidation numbers in a neutral compound is zero. The sum of all oxidation numbers in a polyatomic (many-atom) ion is equal to the charge on the ion.

This rule often allows chemists to calculate the oxidation number of an atom that may have multiple oxidation states, if the other atoms in the ion have known oxidation numbers.

Rule 4: The oxidation number of an alkali metal (IA family) in a compound is +1; the oxidation number of an alkaline earth metal (IIA family) in a compound is +2.

Rule 5: The oxidation number of oxygen in a compound is usually –2. If, however, the oxygen is in a class of compounds called peroxides.

(for example, hydrogen peroxide), then the oxygen has an oxidation number of –1. If the oxygen is bonded to fluorine, the number is +1.

Rule 6: The oxidation state of hydrogen in a compound is usually +1. If the hydrogen is part of a binary metal hydride (compound of hydrogen and some metal like LiH, NaH, CaH_2), then the oxidation state of hydrogen is –1.

Rule 7: The oxidation number of fluorine is always –1. Chlorine, bromine, and iodine usually have an oxidation number of –1, unless they're in combination with an oxygen or fluorine.

TASK: Assign ON to all the elements in the following compounds and ion:

(i) Li_2O, (ii) HNO_3, (iii) $Cr_2O_7$2- (iv) MnO_4-

- **Disproportionation** is a chemical reaction, typically a redox reaction, where a molecule is transformed into two or more dissimilar products. In a redox reaction, the species is simultaneously oxidized and reduced to form at least two different products.

- Consider the following skeleton reaction which takes place in *acidic medium*:

$$\overset{+7}{MnO_4^-}(aq) + \overset{+2}{Fe^{2+}}(aq) \rightarrow \overset{+2}{Mn^{2+}}(aq) + \overset{+3}{Fe^{3+}}(aq)$$

- In this reaction MnO_4^- (purple color) acts as an oxidizing agent in acidic solution and itself is reduced to Mn^{2+} (pale pink to colorless). Iron (II) (pale green) is oxidized to Fe^{3+} (pale yellow to colorless).

Oxidising agents

- are electron rich species
- oxidized other but reduced itself

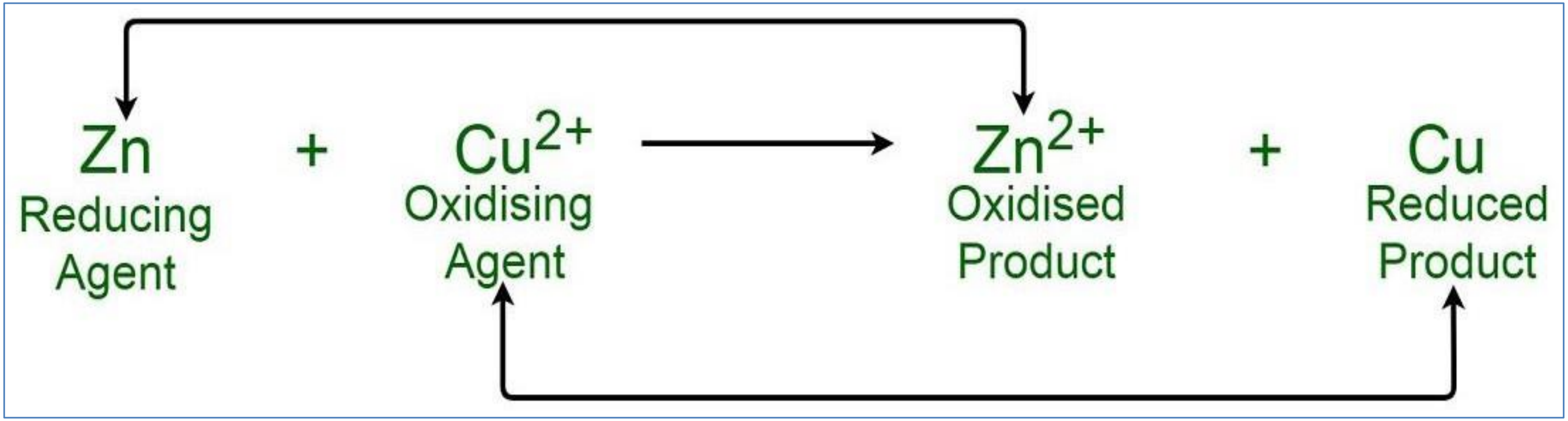
Example: All non- metals (except inert gas), F_2, Cl_2, Br_2, I_2, O_2, HNO_3, Conc.H_2SO_4, $KMnO_4$, $K_2Cr_2O_7$, etc

$$Cl_2\ (g) \rightarrow 2Cl\ (g)\ Cl\ (g) + e^- \rightarrow Cl^-(g)$$

Reducing agents

- are electron-deficient species
- oxidized itself but reduced other
- **Example:** All metals, Hydrogen, carbon, CO, H_2S, SO_2, Oxalic acid ($H_2C_2O_4$), $SnCl_2$.

Oxidizing and Reducing agents in a Reaction:



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