SAN JACINTO COLLEGE

STUDENT SUCCESS CENTER

Oxidation- Reduction (Redox) Reactions

Introduction

Oxidation-reduction, or redox, reactions are important types of reactions. A redox reaction involves electron transfer between two chemical species. In other words, the important aspect in these reactions is the transfer of electrons from one reactant to the other. Formation of a compound from its elements, decomposition into simple elements, and combustion reactions are considered among redox reactions. Oxidation states of an element identify the presence of a redox reaction, since the oxidation number of a molecule, atom, or ion changes by gaining or losing an electron. The following are common types of redox reactions:

Combination:	$A + B \rightarrow AB$
Decomposition:	$AB \rightarrow A + B$
Single Replacement:	$A + BC \rightarrow AC + B$
Double Replacement:	$AB + CD \rightarrow AD + CB$

Oxidation States

In essence, the oxidation state/number (O.N.) of an element is defined as the number of electrons that an atom gains or loses in ionic compounds or shares in molecular compounds. Movement of electron charges could happen in both ionic or molecular compounds. The electrons move from cation to anion in the formation of an ionic compound to perform a complete transfer; however, there is an overall shift in electron charge rather than a transfer in molecular (covalent) compounds.

There are several general rules in determining the oxidation states:

- The O.N. of an atom in its elemental form is 0.
- The O.N. of an ion is equal to its ionic charge.
- The sum of O.N. values in a molecular or ionic compound is equal to zero.
- The sum of O.N. values for atoms in a polyatomic ion is equal to the charge of ion.

Table 1 on the next page shows the specific rules in determining the oxidation numbers of periodic table atoms [1] [2].

Elements of Redox Reactions

In redox reactions, normally one reactant is oxidized and one reactant is reduced at the same time. An atom is oxidized if its O.N. increases while an atom is reduced of its O.N. decreases. **Oxidation** is defined as a loss of electrons in the process, and **reduction** is the gain of electrons. The element that is being reduced is called the **oxidizing agent** since it causes the oxidation of another element by gaining electrons. In addition, the oxidized element is known as a **reducing agent** when it reduces the other element by losing electrons.

Atoms	Oxidation Numbers (ON)
Group 1A (1)	+1
Group 2A (2)	+2
Hydrogen	+1 with nonmetals -1 with metals and Boron
Fluorine	-1
Oxygen	-1 in peroxides -2 in others (except F)
Group 7A (17)	-1 (except O)

Table 1 Rules on assigning oxidation numbers

Redox reactions can be considered to have two hypothetical half-reactions: Oxidation and reduction half reactions. The sum of both half reactions maintains the charge neutrality of the original reaction. The following examples show a typical redox reaction with its elements [1].

Main reaction: $Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$
Oxidation half-reaction: $Cu(s) \rightarrow Cu^{2+}$
Reduction half-reaction: $2Ag^{+}(aq) \rightarrow 2Ag(s)$

Balancing Redox Reactions

The **Half Equation Method** is used in balancing redox reactions. In this method, the equation is divided into two oxidation and reduction half reactions. Each equation is balanced by adjusting coefficients independently, by adding electrons, H^+, H_2O, and OH^- (Basic conditions) [3]. The following shows the step-by-step method to balance the reaction below in basic conditions:

$$MnO_4^{-1}(aq) + I^{-}(aq) \rightarrow Mn^{2+}(aq) + I_2(s)$$

1. Write the half reactions.

   Reduction: $MnO_4^{-1}(aq) \rightarrow Mn^{2+}(aq)$

   Oxidation: $I^{-}(aq) \rightarrow I_2(s)$

2. Balance elements other than O and H.

   Reduction: $MnO_4^{-1}(aq) \rightarrow Mn^{2+}(aq)$

   Oxidation: $2I^{-}(aq) \rightarrow I_2(s)$

3. Balance the O and H atoms by adding H_2O molecules and H^+ ions respectively.

   Reduction: $MnO_4^{-1}(aq) + 8H^{+}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

Oxidation: $2I^{-}(aq) \rightarrow I_2(s)$

4. Sum the charges on each side for each half reactions. Add electrons (e-) to balance charges.

   Reduction: $MnO_4^{-1}(aq) + 8H^{+}(aq) + 5e^{-} \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

   Oxidation: $2I^{-}(aq) \rightarrow I_2(s) + 2e^{-}$

5. The e- on each side should be equal; if not, multiply appropriate integers to half reactions to make them equal.

   Reduction: $2 * \{MnO_4^{-1}(aq) + 8H^{+}(aq) + 5e^{-} \rightarrow Mn^{2+}(aq) + 4H_2O(l)\}$

   $2MnO_4^{-1}(aq) + 16H^{+}(aq) + 10e^{-} \rightarrow 2Mn^{2+}(aq) + 8H_2O(l)$

   Oxidation: $5 * \{2I^{-}(aq) \rightarrow I_2(s) + 2e^{-}\}$

   $10I^{-}(aq) \rightarrow 5I_2(s) + 10e^{-}$

6. Add the half reactions together, and cancel out the electrons.

$$2MnO_4^{-1}(aq) + 16H^{+}(aq) + \cancel{10e^{-}} \rightarrow 2Mn^{2+}(aq) + 8H_2O(l)$$

$$10I^{-}(aq) \rightarrow 5I_2(s) + \cancel{10e^{-}}$$

$$2MnO_4^{-1}(aq) + 10I^{-}(aq) + 16H^{+}(aq) \rightarrow 2Mn^{2+}(aq) + +5I_2(s) + \ 8H_2O(l)$$

7. If the reaction is in basic condition, add OH^{-} ions to neutralize H^{+} ions.

$$2MnO_4^{-1}(aq) + 10I^{-}(aq) + 16H^{+}(aq) + 16OH^{-}(aq) \rightarrow 2Mn^{2+}(aq) + +5I_2(s) + \ 8H_2O(l) + 16OH^{-}(aq)$$

8. Balance the produced H_2O molecules in the last step.

$$2MnO_4^{-1}(aq) + 10I^{-}(aq) + 16H_2O(aq) \rightarrow 2Mn^{2+}(aq) + +5I_2(s) + \ 8\cancel{H_2O}(l) + 16OH^{-}(aq)$$

***Final** :* $2MnO_4^{-1}(aq) + 10I^{-}(aq) + 8H_2O(aq) \rightarrow 2Mn^{2+}(aq) + +5I_2(s) + 16OH^{-}(aq)$

Works Cited

1. M. S. Silberberg, Principles of General Chemistry, New York: McGraw-Hill, 2007.
2. N. J. Tro, Chemistry; A Molecular Approach, Upper Saddle River, NJ: Pearson, 2014.
3. "Chemistry LibreTexts," Libretexts, 27 April 2018. [Online]. Available: https://chem.libretexts.org/Core/Analytical_Chemistry/Electrochemistry/Redox_Chemistry/Balancing_Redox_reactions