

Introduction to Oxidation-Reduction (Redox) Reactions

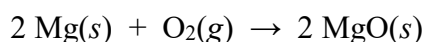
In Unit 4, you have seen examples of the following reaction types.

precipitation reaction: Two aqueous solutions of ionic compounds are combined. The cations of one solution combine with the anions of another solution to form an insoluble ionic compound, known as a precipitate.

acid-base reaction: A Bronsted-Lowry acid (H^+ donor) reacts with a Bronsted-Lowry base (H^+ acceptor). The H^+ is transferred from the acid to the base.

We will now study another type of reaction known as oxidation-reduction or redox.

Consider the following balanced chemical equation for the reaction between magnesium (Mg) and oxygen (O_2).



This equation can be split up into two separate processes.

| | |
|---|--|
| An atom of Mg changes into a Mg^{2+} ion | A molecule of O_2 changes into two O^{2-} ions |
| $\text{Mg} \rightarrow \text{Mg}^{2+}$ | $\text{O}_2 \rightarrow 2 \text{O}^{2-}$ |

You should know and understand each of the following definitions about oxidation and reduction.

oxidation: a substance loses electrons

reduction: a substance gains electrons

You can use the mnemonic device **OIL RIG** to help you remember these definitions.

OIL: Oxidation is the Loss of electrons

RIG: Reduction is the Gain of electrons

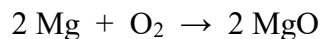
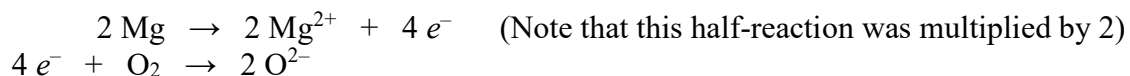
Equations that show either oxidation or reduction separately are known as **half-reactions**.

- In an oxidation half-reaction, electrons appear on the right side of the equation.
- In a reduction half-reaction, electrons appear on the left side of the equation.

The half-reactions for the reaction involving Mg and O_2 are written as follows. Notice that the overall charge is balanced on both sides of the equation.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|--|--|
| $\text{Mg} \rightarrow \text{Mg}^{2+} + 2 e^-$ | $4 e^- + \text{O}_2 \rightarrow 2 \text{O}^{2-}$ |

- The number of electrons lost in the oxidation half-reaction must be equal to the number of electrons gained in the reduction half-reaction.
- One (or both) of the half-reactions may need to be multiplied by a certain number so that the electrons can be cancelled out on both sides of the equation.
- When the two half-reactions are added together, you should get the overall balanced redox equation.



(Electrons do not appear in the overall redox equation.)

In each of the following, you are given two separate half-reactions. You need to add the two half-reactions together to produce a balanced redox equation based on the rules listed on the previous page.

1.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|--|---|
| $\text{Zn} \rightarrow \text{Zn}^{2+} + 2 e^{-}$ | $2 e^{-} + 2 \text{H}^{+} \rightarrow \text{H}_2$ |

2.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|--|--|
| $\text{Al} \rightarrow \text{Al}^{3+} + 3 e^{-}$ | $2 e^{-} + \text{Cu}^{2+} \rightarrow \text{Cu}$ |

3.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|---|---|
| $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^{-}$ | $5 e^{-} + \text{MnO}_4^{-} + 8 \text{H}^{+} \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O}$ |

4.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|---|--|
| $2 \text{Cl}^{-} \rightarrow \text{Cl}_2 + 2 e^{-}$ | $6 e^{-} + 14 \text{H}^{+} + \text{Cr}_2\text{O}_7^{2-} \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}$ |

5.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|--|---|
| $\text{Ni} \rightarrow \text{Ni}^{2+} + 2 e^{-}$ | $3 e^{-} + 4 \text{H}^{+} + \text{NO}_3^{-} \rightarrow \text{NO} + 2 \text{H}_2\text{O}$ |

Oxidation Numbers

In an oxidation-reduction (redox) reaction, electrons are transferred from the substance that is oxidized to the substance that is reduced. One of the ways to recognize that a reaction is a redox reaction is to look for changes in the oxidation numbers of the elements involved.

Oxidation numbers are a bookkeeping system that is used to keep track of any changes in electrons being gained or lost. The following rules are used when assigning oxidation numbers to the atoms in a molecule or an ion.

| Rule | Examples |
|--|--|
| #1: For an atom in its elemental form, the oxidation number is zero. | $\text{H}_2 = \text{zero}$ $\text{Li} = \text{zero}$ $\text{N}_2 = \text{zero}$ |
| #2: For a monoatomic ion, the oxidation number is the charge on the ion. | $\text{Na}^+ = +1$ $\text{Mg}^{2+} = +2$ $\text{Cl}^- = -1$ |
| #3: The oxidation number of Group 1 metals is +1 in all compounds, and the oxidation number of Group 2 metals is +2 in all compounds. | $\text{Na} = +1$ in compounds such as NaCl , NaNO_3 , and Na_2O . $\text{Mg} = +2$ in compounds such as MgBr_2 , MgSO_4 , and MgS . |
| #4: The oxidation number of F is -1 in all compounds. | $\text{F} = -1$ in compounds such as NaF , CH_3F , and HF . |
| #5: The oxidation number of H is usually +1 when bonded to nonmetals and -1 when bonded to metals. | $\text{H} = +1$ in compounds such as H_2O , NH_4Cl , and CH_4 . $\text{H} = -1$ in compounds such as LiH , NaH , and CaH_2 . |
| #6: The oxidation number of O is usually -2 in most compounds. Two important exceptions are H_2O_2 and OF_2 . | $\text{O} = -2$ in compounds such as CO_2 , KNO_3 , and H_2SO_4 . $\text{O} = -1$ in H_2O_2 $\text{O} = +2$ in OF_2 |
| #7: The oxidation number of Cl, Br, and I is -1 in most binary compounds. However, the oxidation number of Cl, Br, and I has a positive value when bonded with O or F. | $\text{Cl} = -1$ in compounds such as KCl , CCl_4 , and SCl_2 . $\text{Cl} = +1$ in NaClO and HClO $\text{Cl} = +3$ in NaClO_2 and HClO_2 $\text{Cl} = +5$ in NaClO_3 and HClO_3 $\text{Cl} = +7$ in NaClO_4 and HClO_4 |

- The sum of the oxidation numbers of all atoms in a neutral compound is equal to zero.
- The sum of the oxidation numbers of all atoms in a polyatomic ion is equal to the overall charge on the polyatomic ion.

6. Based on the rules on the previous page, assign oxidation numbers to each atom in each substance.

| Substance | Oxidation Numbers | | |
|--|-------------------|------|-----|
| Fe | Fe = | | |
| Cl ₂ | Cl = | | |
| P ₄ | P = | | |
| NaBr | Na = | Br = | |
| Ca ₃ N ₂ | Ca = | N = | |
| Fe ₂ O ₃ | Fe = | O = | |
| MgS | Mg = | S = | |
| SO ₂ | S = | O = | |
| CH ₃ F | C = | H = | F = |
| CHF ₃ | C = | H = | F = |
| CO | C = | O = | |
| OF ₂ | O = | F = | |
| H ₂ O ₂ | H = | O = | |
| KH | K = | H = | |
| CaH ₂ | Ca = | H = | |
| HBr | H = | Br = | |
| HBrO | H = | Br = | O = |
| HBrO ₂ | H = | Br = | O = |
| HBrO ₃ | H = | Br = | O = |
| Na ₂ CO ₃ | Na = | C = | O = |
| BaCrO ₄ | Ba = | Cr = | O = |
| NH ₄ ⁺ | N = | H = | |
| OH ⁻ | O = | H = | |
| NO ₃ ⁻ | N = | O = | |
| ClO ₄ ⁻ | Cl = | O = | |
| SO ₃ ²⁻ | S = | O = | |
| SO ₄ ²⁻ | S = | O = | |
| Cr ₂ O ₇ ²⁻ | Cr = | O = | |
| PO ₄ ³⁻ | P = | O = | |
| H ₃ O ⁺ | H = | O = | |

7. Each of the following equations represents an oxidation-reduction reaction.
- Assign oxidation numbers to each element on each side of the equation.
 - Determine which element is oxidized and which element is reduced in the reaction.
 - If the oxidation number increases, the element is oxidized.
 - If the oxidation number decreases, the element is reduced.

| Balanced Equation | Element that is Oxidized | Element that is Reduced |
|---|--------------------------|-------------------------|
| $\text{N}_2 + 3 \text{H}_2 \rightarrow 2 \text{NH}_3$ | | |
| N = H = | N = H = | |

| Balanced Equation | Element that is Oxidized | Element that is Reduced |
|---|----------------------------|-------------------------|
| $3 \text{Fe}(\text{NO}_3)_2 + 2 \text{Al} \rightarrow 3 \text{Fe} + 2 \text{Al}(\text{NO}_3)_3$ | | |
| Fe = N = O = Al = | Fe = N = O = Al = | |

| Balanced Equation | Element that is Oxidized | Element that is Reduced |
|---|---------------------------|-------------------------|
| $2 \text{BrO}_3^- + 3 \text{N}_2\text{H}_4 \rightarrow 2 \text{Br}^- + 6 \text{H}_2\text{O} + 3 \text{N}_2$ | | |
| Br = O = N = H = | Br = O = N = H = | |

| Balanced Equation | Element that is Oxidized | Element that is Reduced |
|--|---------------------------|-------------------------|
| $\text{P}_4 + 10 \text{HClO} + 6 \text{H}_2\text{O} \rightarrow 4 \text{H}_3\text{PO}_4 + 10 \text{HCl}$ | | |
| P = H = Cl = O = | P = H = Cl = O = | |

| Balanced Equation | Element that is Oxidized | Element that is Reduced |
|--|--------------------------|-------------------------|
| $2 \text{C}_2\text{H}_6 + 7 \text{O}_2 \rightarrow 4 \text{CO}_2 + 6 \text{H}_2\text{O}$ | | |
| C = H = O = | C = H = O = | |

| Balanced Equation | Element that is Oxidized | Element that is Reduced |
|--|---------------------------|-------------------------|
| $\text{Mn} + \text{H}_2\text{SO}_4 \rightarrow \text{MnSO}_4 + \text{H}_2$ | | |
| Mn = H = S = O = | Mn = H = S = O = | |

8. Classify each of the following reactions as precipitation, acid-base, or redox.

| Balanced Equation for the Reaction | Reaction Type |
|--|---------------|
| $\text{Zn}(s) + 2 \text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)$ | |
| $\text{Ca}(\text{OH})_2(aq) + 2 \text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + 2 \text{H}_2\text{O}(l)$ | |
| $\text{AgNO}_3(aq) + \text{HCl}(aq) \rightarrow \text{AgCl}(s) + \text{HNO}_3(aq)$ | |
| $2 \text{KOH}(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow 2 \text{H}_2\text{O}(l) + \text{K}_2\text{SO}_4(aq)$ | |
| $\text{Pb}(\text{NO}_3)_2(aq) + \text{K}_2\text{SO}_4(aq) \rightarrow 2 \text{KNO}_3(aq) + \text{PbSO}_4(s)$ | |
| $\text{PbS}(s) + 4 \text{H}_2\text{O}_2(aq) \rightarrow 4 \text{H}_2\text{O}(l) + \text{PbSO}_4(s)$ | |

Net Ionic Equations

9. For each of the following redox reactions, write the balanced net ionic equation.

