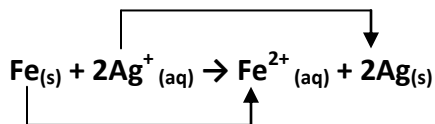


Oxidation and Reduction

Oxidation and reduction reactions involve the transfer of electrons from one atom to another – one atom losing electrons and another atom gaining electrons. This type of reaction is often called a “redox” reaction (short for reduction/oxidation) as a reminder that both processes are going on at the same time.

For the reaction:

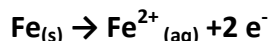


- The reduction half-reaction occurs when electrons are gained or added, causing an atom to become more negative (less positive):



In order for a reduction reaction to occur, another atom must be oxidized – in other words, if one atom is going to gain electrons, another atom has to give electrons away. The atom that is reduced is also known as the “oxidizing agent” because it is causing oxidation of another element.

- The oxidation half-reaction occurs when electrons are lost or donated, causing an atom to become less negative (more positive):



Again, if oxidation is to occur, another atom must be reduced – if one atom is going to give electrons away, another atom must accept them. The atom that is oxidized is also called the “reducing agent” because it is causing reduction of another element.

Notice that the number of electrons gained in the reduction reaction is exactly the same as the number of electrons that were given away in the oxidation reaction. In order for a redox equation to balance, the same number of electrons must be transferred in each half-reaction.

- To help identify oxidation and reduction, remember:

OILRIG: Oxidation Is Losing Electrons, Reduction Is Gaining Electrons

or

LEO the Lion says **GER:** Losing Electrons is Oxidation, Gaining Electrons is Reduction.

Assigning Oxidation Numbers

Oxidation numbers provide a way to identify the number of electrons gained or lost by an element. The following rules help identify the oxidation number of an element:

- Elements in their natural (uncombined) state** have an oxidation number of 0 (zero).
- The oxidation number of any **monatomic ion** (only 1 element) is equal to its ionic charge.
- The oxidation number of **hydrogen** is usually +1.
 - Exception: In metal hydrides (such as NaH), the oxidation number of hydrogen is -1.
- The oxidation number of **oxygen** is usually -2.
 - Exception: In peroxides (such as H₂O₂), the oxidation number of oxygen is -1.
- The **sum of the oxidation numbers in a compound** equals 0 (zero).
- The **sum of the oxidation numbers of a polyatomic ion** is equal to the charge of the ion.

Balancing Redox Equations

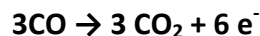
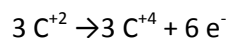
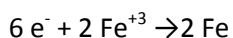
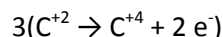
I. Oxidation Number Method

1. Assign oxidation numbers to every element in the equation.
2. Identify elements that have a change in oxidation number.
3. Show the half-reactions for oxidation and reduction, identifying the number of electrons gained or lost.
4. Use coefficients to make the number of electrons gained the same as the number of electrons lost.
5. Bring the coefficients into the equation by balancing each element.
6. Balance any remaining elements by inspection.

Example:



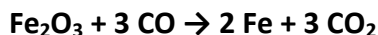
The lowest common multiple of 2 and 3 is 6, so there must be 6 electrons gained and 6 electrons lost.



(There should be 2 Fe atoms on each side of the equation – Fe₂O₃ already has 2 Fe, but Fe has only 1, so it needs a coefficient of 2)

(There should be 3 C atoms on each side of the equation, so both CO and CO₂ need a coefficient of 3)

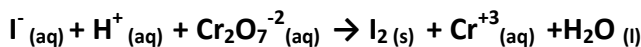
Add the 2 half-reactions and verify that the result is balanced:



II. Half-reaction Method

1. Identify the 2 half-reactions by inspection or by assigning oxidation numbers.
2. For each half-reaction, balance any elements other than hydrogen and oxygen.
3. Balance oxygen by adding H₂O's to the side that is oxygen-poor.
4. Balance hydrogen by adding H⁺ ions to the side that is hydrogen-poor.
5. Balance overall charges by adding electrons to the more positive side of the half-reaction.
6. Use coefficients to make the number of electrons gained in the reduction half-reaction the same as the number lost in the oxidation half-reaction.
7. Add the two half-reactions together, canceling out as needed.

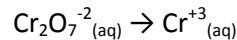
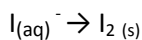
Example:



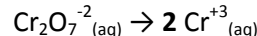
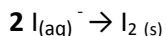
(Oxidation half-reaction)

(Reduction half-reaction)

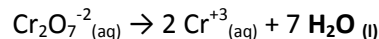
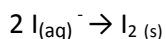
Half-reactions:



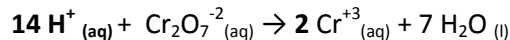
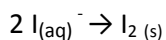
Balance elements (except H and O)



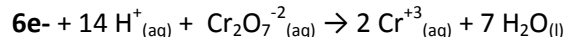
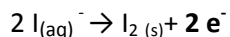
Add H₂O to balance O's



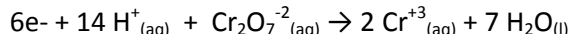
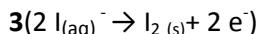
Add H⁺ to balance H's



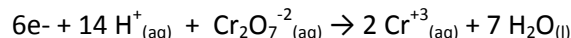
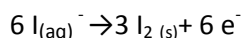
Add electrons to balance charges



Use coefficients to match # of e⁻



Add half-reactions



Balanced equation:

