Oxidation and Reduction

Outlines, explanations, examples, and answer keys

What is a “Redox” Reaction?

Redox - Reduction/Oxidation. Effectively, these are reactions in which the oxidation number of a species changes. Simply put, one or more electrons (e^-) transferred from one atom in the reaction to another.

Oxidation Numbers

Oxidation numbers are a tool used by chemists to quickly determine the number of electrons in a given atom. It can also be interpreted as the “charge” of the atom. This is done by determining the deviation in the number of electrons from an expected value.

If Protons have a charge of +1 and Electrons have a charge of -1, in an atom where # of Protons = # of Electrons, then the oxidation number (and charge) is 0. This is the case for all pure elements (not bonded to any other atoms of differing elements). Ag is pure, AgCl is not!

Because Electrons are the source of chemical bonding, their abundance and location is what oxidation numbers track

Rules to determine Oxidation Numbers

1. The oxidation number of any pure element is zero. This includes elements that form a pair in their pure form such as H2 or N2
2. The oxidation number of a monoatomic ion is equal to its charge. Cl- has a oxidation number of -1
3. When calculating the oxidation numbers of elements that form a compound, remember that their sum of oxidation numbers is equal to the charge of the total compound. Hydroxide (OH-) has a total charge of -1. It can be inferred that the Oxygen has a charge of -2. While the Hydrogen has a charge of +1. (1 - 2 = -1)
    a. Generally speaking, you can predict oxidation numbers of individual elements in a molecule provided they form monoatomic ions of known charges

Electrons and Oxidation Numbers

Adding a singular electron lowers an atom's oxidation number by 1, as an electron has a charge of negative 1. Removing an electron increases the oxidation number by 1. This can be visualized in the next slide.

Remember, the oxidation number is NOT the number of electrons added or removed. However, you can think of it as the inverse of this value.

Half - Reactions

Similar to balancing an equation, half-reactions are obtained from balanced oxidation/reduction reactions

They are split into the Oxidizing Half-rxn, and the Reducing Half-rxn. The key change is that electrons (e^-) are shown as a component of the reaction. For instance:

$$Mn^{2+} + 2e^- \rightarrow Mn_{(s)}$$

In this reaction, Manganese is being **reduced** by two electrons to form elemental Manganese. These electrons are now part of the electron shell, and their additional negative charge results in a net zero charge for the atom.

Final Vocabulary Check

Oxidized - Element/molecule lost electrons (more positive charge)

Reduced- Element/molecule gained electrons (more negative charge)

Oxidizing agent - Oxidizes another compound, being reduced in the process

Reducing agent- Reduces another compound, becoming oxidized in the process

For the following examples:

Balance the following redox reactions. In each case

a. Identify the oxidation numbers for each element in the reaction
b. give the balanced half-reactions; identify the oxidation half-reaction and the reduction half-reaction.
c. give the balanced net reaction.
d. identify the oxidizing agent and the reducing agent.

$$Al_{(s)} + Fe_2O_{3(s)} \rightarrow Fe_{(l)} + Al_2O_{3(l)}$$

$$Al_{(s)} + Fe_2O_{3(s)} \rightarrow Fe_{(l)} + Al_2O_{3(l)}$$

2Al(s) + Fe2O3(s) → 2Fe(l) + Al2O3(l)

0 +3 -2 0 -2+3

2Al(0) → 2Al(+3) + 6e- (Oxidation)

2Fe(+3) + 6e- → 2Fe(0) (Reduction)

Al is the Reducing Agent. Fe is the Oxidizing agent. Iron Oxide Oxidized Aluminum metal.