Introduction

Oxidation-reduction reactions (usually shortened to **redox** reactions) are a very common kind of reaction. Many of the chemical reactions you have seen outside of class are redox reactions. For example, the process of **rusting** (formation of iron(III) oxide from iron) is a redox reaction. Most **combustions** are also redox reactions.

Redox reactions can't be treated as simple exchange reactions like the precipitations and most neutralizations.

Definition

What **is** a redox reaction? A redox reaction actually involves two of what we term **half-reactions**. These half-reactions are called **oxidation** and **reduction**. Let's look at rusting for an illustration of what oxidation and reduction are:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

What actually happens in this reaction? Let's first look at the fate of the iron - it changes from solid metallic iron to iron(III) ion (in the ionic compound Fe_2O_3). In equation form:

$$Fe(s) \rightarrow Fe^{3+} + 3\ e^-$$

The iron had to **lose electrons** (the 3 e^-) to go from metallic iron (zero charge) to iron (III) (+3 charge). This **loss of electrons** is called **oxidation**, and the iron is said to have been **oxidized**.

You may be wondering where the electrons released by the iron go. Let's find out by looking at the **other** reactant - oxygen.

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

Oxygen starts out as molecular oxygen (with zero charge). Oxygen ends up as the oxide ion, which has a -2 charge. It gets the negative charge by **accepting electrons**, and this gain of electrons is called **reduction**.

*Hint: You can think of "reduction" as meaning a **reduction in charge** (which can only be caused by a **gain of electrons**)*

If you add up the two **half-reactions**, oxidation and reduction, you have – after balancing - the overall **redox reaction** for the rusting of iron.

Related terms - oxidizing and reducing agents

Now that you know what oxidation and reduction are, we need to define two important terms related to redox reactions. They're important because they're often seen on warning labels on reagents in the lab and in industry. Occasionally, you might find one of these labels on a chemical bottle around the home.

An **oxidizing agent**, or **oxidizer**, is a chemical that causes **something else to be oxidized**. The oxidizing agent itself is reduced in the chemical reaction. Oxidizing agents, then, crave electrons (What is the oxidizing agent in rusting?). Of the simple chemical species, you'd find oxidizing agents among the elements that readily form **anions**.

A **reducing** agent, on the other hand, causes **something else to be reduced**. The reducing agent is oxidized in the chemical reaction. Reducing agents have electrons to spare - that is, they readily give up electrons (What is the reducing agent in rusting?). Of the simple chemical species, you'd find reducing agents among the elements that readily form **cations**.

If you see a bottle of a **strong oxidizing agent** (there are more common strong oxidizing agents than strong reducing agents), you'd be well advised to keep it away from combustible material. Redox reactions involving strong oxidizing agents can release lots of energy in the form of heat, which can cause a fire.

Types of redox reactions

There are so many different reactions that can be classified as redox reactions that we actually subdivide them into smaller categories. If you've taken high school chemistry, you're probably familiar with some of these categories. We will discuss four of these.

1. **Combination reactions** involve two or more substances **combining** to form a **single substance**. These reactions are sometimes called synthesis reactions, because you are synthesizing a new chemical from several less complex starting materials.

General form of a combination reaction
$A + B + ... \rightarrow D$
Example
$2K(s) + F_2(g) \rightarrow 2KF(s)$ (K is **oxidized**, F is **reduced**)

2. **Combustion reactions** involve a substance reacting **rapidly** with O_2 (molecular oxygen) to produce heat and **flame**. Some of these can be classified as **combinations** as well. The products of a combustion reaction are **oxides** – combinations of a single element with oxygen. Oxygen is the oxidizing agent in combustion reactions.

General form of a combustion reaction
$A + O_2 \rightarrow AO$
Example
$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$ (Mg is **oxidized**, O is **reduced**)

3. **Decomposition reactions** involve a substance **breaking down** into several other simpler substances. In essence, these reactions are the opposite of combination reactions.

General form of a decomposition reaction
$A \rightarrow B + C + ...$
Example
$2H_2O_2 \rightarrow 2H_2O + O_2$

4. **Single displacement reactions** occur when one element **replaces** or "kicks out" another in a compound. These reactions **usually** involve metal species, and a more **active** metal will replace a less active metal in a compound. The element by itself at the beginning of the reaction is oxidized, while the element it replaces is reduced.

General form of a single displacement reaction
$A + BC \rightarrow AC + B$
Example
$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$ (Cu has been **oxidized**, and Ag has been **reduced**!)

In the example, the copper (a **more active** metal) has replaced silver (a **less active** metal) in silver nitrate, forming copper(II) nitrate and silver metal. **Activity** is essentially a measure of how readily an element will give up an electron. How do you tell whether one metal or another is more active? Use an activity series, which is a compilation of experimental data from single replacement reactions.

Activity Series (shortened)

Very active metals →	Li
(These will displace hydrogen	K
in acids and - though less quickly - in water)	Na
	Mg
Active metals →	Zn
	Fe
	Pb
Hydrogen (in acids) →	H
	Cu
Relatively inactive metals →	***Ag***

The reaction of metallic sodium with water is also a single displacement reaction:

$$\mathbf{Na(s)} + \mathbf{H}OH(l) \rightarrow \mathbf{Na}OH(aq) + \mathbf{H_2(g)}$$

The sodium displaces hydrogen from water to form the hydroxide and hydrogen gas as products! (*H_2O is written as HOH here to illustrate the displacement more clearly.*)

Oxidation states

Oxidation states are used to show electron transfer in redox reactions. These oxidation states are a pseudo-charge used to account for electrons in compounds where simply looking at charges on ions is not sufficient (in other words, **molecular compounds**).

For simple monatomic ions, oxidation state is the same thing as the charge on the ion. For molecular species, oxidation states are assigned by **rules**. We will deal mainly with ionic compounds in our discussions.

Summary

In this note pack, we discussed redox (oxidation-reduction) reactions and the different types of redox reaction that we encounter in the lab. You should know that simple redox reactions involve electron transfer between reagents. You should be familiar with the terms "oxidation", "reduction", "oxidizing agent", and "reducing agent". We also discussed the four common types of redox reaction. You should know the general form of each reaction type and be able to tell, given a reaction, what type it is.