Learning Objectives

- Use given guidelines to determine oxidation numbers of atoms in a compound.
- Differentiate between oxidation numbers and charges
- Use oxidation number to determine which species in a reaction is oxidized and which is reduced, which is the oxidizing agent and reducing agent.
- Recognize and identify redox reactions
- Determine if single displacement redox reactions are spontaneous or not based on activity series.

Introduction

Another way of classifying reactions is do they involve the transfer of electrons from one chemical species to another? If so, it is an oxidation/reduction reaction or Redox Reaction for short. The definitions for oxidation and reduction are:

**Oxidation:** the loss of electrons  
**Reduction:** the gain of electrons

So far we have covered two types of reactions that must involve oxidation and reduction, the single displacement reaction and the combustion reaction. The single displacement reaction is easy to see as it involves neutral metals becoming ions. The combustion reaction is more complicated as elemental oxygen is forming covalent bonds and there are no ions to help us identify which species is gaining and which species is losing electrons, and so we will need to introduce the concept of oxidation numbers to tackle this.

Let’s start by looking at a very simple reaction, that of iron and oxygen to form Iron(III) oxide.

\[
2\text{Fe(s)} + 3\text{O}_2(g) \rightarrow 2\text{Fe}^{+3} + 3\text{O}^{-2} = \text{Fe}_2\text{O}_3(s)
\]

In this reaction the iron is getting oxidized because each iron atom loses three electrons and the oxygen is getting reduced because each oxygen gains two electrons. We call oxygen the oxidant (oxidizing agent) because it causes oxidation while becoming reduced (3O₂ +6e⁻ -> 3O²⁻), that is, it must grab the 6 electrons from the iron in order to get reduced. Iron is called the reductant (reducing agent) because it reduces the oxygen while becoming oxidized (2Fe -> 2Fe⁺³ + 6e⁻). That is, it gives 6 electrons to the oxygen causing reduction (as it gets oxidized).

We will start with single displacement reactions because they are the easiest to understand.

Single Displacement Reactions

Single displacement reactions were introduced in section 3.2.1.3 and there were two types, metal single displacement reactions and nonmetal single displacement reactions. Metals exist in two common forms, as an elemental (pure) metal
For as a cation in an ionic compound (FeS). In a metal single displacement reaction a pure metal displaces the cation of another metal in an ionic compound. Nonmetals also exist as pure elements (Cl₂) and as anions in ionic bonds (the also commonly exist in many types of covalent compounds). In a nonmetal single displacement reaction, a pure nonmetal displaces the anion of another nonmetal in an ionic compound. The more active a metal or nonmetal, the more it wants to form an ion, the less active, the more it tends to exist in its elemental form.

**Metal Displacement Reactions**

The more active a metal, the more it wants to exist as an ion. The less active it is, the more it wants to stay as a pure element.

Let's look at the reaction for Zinc metal with Copper(II) ions (in a CuSO₄ salt).

\[
\text{Zn(s) + CuSO}_4{\text{(aq) \rightarrow ZnSO}_4{\text{(aq) + Cu(s)}}}\text{ (Spontaneous Reaction, is actually used in batteries!)}
\]

and compare that to the reaction of Copper metal with Zinc ions (in a ZnCl₂ salt).

\[
\text{Cu(s) + ZnSO}_4{\text{(aq) \rightarrow ZnSO}_4{\text{(aq) + Cu(s)}}}\text{ (No Reaction, nothing happens!)}
\]

So Zinc is a more active metal than Copper. The following video shows how an activity series can be developed that tells you which metals are the most reactive, and which are the least reactive.
Activity Series

Table  shows the ranking of the activity of metals in aqueous solutions. A single displacement reaction will occur when the metallic form of a more active metal comes into contact with a solution containing cations of the less reactive metal. In table  Zinc is above copper, and so we would expect zinc metal to displace copper ions (eq. 3.6.1), and conversely, since copper is below zinc (less active) we would expect no reaction if copper was placed in a solution of zinc ions (eq. 3.6.2). So the reactivity series can be used to predict if a single displacement reaction will occur or not occur. NOTE: There are two activity series in this table. On the left are those of the metals and on the right the halogens. The higher something is on the table, the more it wants to be an ion.
<table>
<thead>
<tr>
<th>Elements</th>
<th>Reaction Occurring</th>
</tr>
</thead>
<tbody>
<tr>
<td>(\text{Li})</td>
<td>React with cold water, replacing hydrogen.</td>
</tr>
<tr>
<td>(\text{K})</td>
<td></td>
</tr>
<tr>
<td>(\text{Ba})</td>
<td></td>
</tr>
<tr>
<td>(\text{Sr})</td>
<td></td>
</tr>
<tr>
<td>(\text{Ca})</td>
<td></td>
</tr>
<tr>
<td>(\text{Na})</td>
<td></td>
</tr>
<tr>
<td>(\text{Mg})</td>
<td></td>
</tr>
<tr>
<td>(\text{Al})</td>
<td></td>
</tr>
<tr>
<td>(\text{Zn})</td>
<td>React with steam, but not cold water, replacing hydrogen.</td>
</tr>
<tr>
<td>(\text{Cr})</td>
<td></td>
</tr>
<tr>
<td>(\text{Fe})</td>
<td></td>
</tr>
<tr>
<td>(\text{Cd})</td>
<td></td>
</tr>
<tr>
<td>(\text{Co})</td>
<td></td>
</tr>
<tr>
<td>(\text{Ni})</td>
<td>Do not react with water. React with acids, replacing hydrogen.</td>
</tr>
<tr>
<td>(\text{Sn})</td>
<td></td>
</tr>
<tr>
<td>(\text{Pb})</td>
<td></td>
</tr>
<tr>
<td>(\text{H}_2)</td>
<td></td>
</tr>
<tr>
<td>(\text{Cu})</td>
<td>Unreactive with water or acids.</td>
</tr>
<tr>
<td>(\text{Hg})</td>
<td></td>
</tr>
<tr>
<td>(\text{Ag})</td>
<td></td>
</tr>
</tbody>
</table>

Table \(\PageIndex{1}\): Activity Series

Activity of Metals

Activity of Halogens
Activity of Metals

\(\text{Pt}\)

\(\text{Au}\)

Table \(\PageIndex{1}\) Activity series of metals and nonmetals reacting in aqueous solutions.

The above table can be used to predict if a single displacement reaction will occur. Note, the position of hydrogen. Metals that are more active than hydrogen will displace the hydrogen from an acid. The alkali metals and heavier alkaline earths on the very top are so reactive that they will displace hydrogen from water, and so are unstable in normal atmospheric conditions. Also note on the top right are the activity series for the halogens. These follow the same principles as the metals

Using the Activity Series

Let’s take a closer look at eq. 3.6.3, the reaction of Zinc(s) with copper(II)Sulfate. The sulfate is a spectator ion and so the net ionic equation is

\[\text{Zn(s) +Cu^{+2}(aq) \rightarrow Zn^{+2}(aq) + Cu(s)}\]

That is the Zinc loses 2 electrons (oxidation)

\[\text{Zn(s) \rightarrow Zn^{+2} + 2e^{-}}\]

while the copper cation gains two electrons (reduction)

\[\text{Cu^{+2} + 2e^{-} \rightarrow Cu(s)}\]

These must occur concurrently, that is the electrons given by the species being oxidized must be given to the species being reduced and for the activity series states that the species with the higher activity will give electrons to a cation of a species of lower activity.

Since Zinc is above Copper, Zince will give two electrons to Cu\(^{+2}\).

Video \(\PageIndex{2}\) goes over how to use the activity series to determine if a single displacement reaction involving metals is spontaneous or not.
Video \(\PageIndex{2}\): 3'13" YouTube showing how to use the activity series to determine if a single displacement reaction is spontaneous or not (https://youtu.be/_tyUhP1LO14)

As for a non metal, the more active halogens prefer to be anions, and so iodide will react with fluorine, but fluoride will not react with iodine (see exercise below).

Exercise \(\PageIndex{1}\)

Predict the products, determine if a reaction occurs, balance the equation and determine phase of all species.

a. \(\text{Al(s)} + \text{Pb(NO}_3\text{)}_4(\text{aq}) \rightarrow \)

b. \(\text{Sn(s)} + \text{CaSO}_4(\text{aq}) \rightarrow \)

c. \(\text{Hg(s)} + \text{NiSO}_4(\text{aq}) \rightarrow \)

d. \(\text{NaCl(aq)} + \text{F}_2(g) \rightarrow \)

Answer a

\[
4\text{Al(s)} + 3\text{Pb(NO}_3\text{)}_4(\text{aq}) \rightarrow 3\text{Pb(s)} + 4\text{Al(NO}_3\text{)}_3(\text{aq})
\]
Answer b

No reaction, tin is less active than calcium

Answer c

No reaction, mercury is less active than nickle

Answer d

\[ 2\text{NaCl(aq)} + \text{F}_2(g) \rightarrow 2\text{NaF(aq)} + \text{Cl}_2(g) \]

Food for Thought

In Video below, we saw that the further down you go on the periodic table the more reactive the alkali metal with water, and yet according to the activity series (table 3.6.1) the trend in activity (Li>K>Na) does not follow a periodic trend. Does anyone want to venture a guess as to what is going on? (Why does the activity series not follow the reactivity with water trend in the video).
Video \(\PageIndex{2}\): 6'16" YouTube by the Royal Society of Chemistry on the reactivity of alkali metals with water. (https://youtu.be/eaChisV5uR0)

Solution

This is actually an important concept that will come up in general chemistry 2. The "reactivity" or "explosiveness" is a function of kinetics and how fast the reaction proceeds. The activity series describes the final state once the reaction is over, and is not a measure of how fast it occurs. So the trend on the video is not the trend of the activity series, but relates to how fast the reaction between water and different alkali metals proceeds. [Note: this is a thought provoking question, which we may come back and visit when we get to the second semester of this class.]

Other Redox Reactions

In a metal single displacement reaction it is very easy to see which reactant is getting oxidized (the pure metal) and which is getting reduced (the cation), but in many redox reactions it is not so clear.

Consider the combustion of methane, for which there are no cations or anions, but for which covalent bonds are being
broken and formed, and so electrons are being transferred between species.

\[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \]

It is clear that on the left side oxygen is neutral, but has it “gained” or "lost" electron density when it goes from bonds between itself to bonds between carbon or hydrogen? To figure this out we introduce the concept of oxidation state to indicate if an atom in a covalent bond is electron rich or electron poor, and we use oxidation numbers to keep track of this in a reaction and determine if an atom was oxidized or reduced, even when no ion was formed.

---

**Oxidation Numbers**

Oxidation numbers indicate an atom's "charge" state, even if it is not a charged ion. There are three fundamental oxidation states, neutral, positive or negative. If you have an atom that is not bonded to anything, its oxidation state is its charge. If it is covalently bonded, its oxidation state is an indicator of whether electron density has moved toward the atom or away from it.

- [0] Oxidation States indicate element is neutral
- [+ ] Oxidation States indicate element is “electron poor”
- [- ] Oxidation States indicate element is “electron rich”

At this point we have a set of rules for determining Oxidation state. In chapter 7 we will gain a deeper understanding of these rules, and in general chemistry 2 we will learn how to use oxidation numbers when we balance chemical equations for charge.

---

**Rules for Identifying Oxidation Numbers**

The following rules allow you to identify the oxidation state of each atom in a covalent bonded molecule or polyatomic ion. Note, the sum of the oxidation states of all species equals the charge of the compound, being zero if it is a neutral molecule.

1. Oxd # = 0 for pure elements
2. Oxd # = charge of monatomic ion
3. Oxd # of F = -1 in compounds with other elements
4. Oxd # of Cl, Br & I = -1 in compounds except with Oxygen & Fluorine
5. Oxd # of H= +1, except for Metal Hydrides (-1)
6. Oxd # of O = -2, except with fluorides, Peroxides (-1) and Superoxides (-1/2)
7. The Sum of the oxd #’s of all elements in a compound = 0, and = the charge of a polyatomic ion

Video \(\PageIndex{3}\) shows how to calculate the oxidation state for the atom in chlorine and sodium thiosulfate (which is also covered in example \(\PageIndex{1}\)).

Example \(\PageIndex{2}\)

Calculate the oxidation state for each atom in sodium thiosulfate \(\text{Na}_2\text{S}_2\text{O}_3\). (From the above rules you will know the oxidation state of every element except one, and can algebraically solve for that. See Video \(\PageIndex{3}\) if you need additional help).

Solution

\(\text{Na}_2\text{S}_2\text{O}_3\)

Given: \([\text{Na}] = +1\) \([\text{S}] = x\) \([\text{O}] = -2\)

Set up equation, not here the sum equals zero, but if it was a polyatomic ion, it would equal the charge of the ion:

\[(2)[\text{Na}] + (2)[\text{S}] + (3)[\text{O}] = 0\]

Solve Equation for unknown

\[(2)[1] + (2)[x] + (3)[-2] = 0\]

\[2 + 2x - 6 = 0\]

\[2x = 4\]

\[x = 2\]
Exercise \( \PageIndex{9} \)

Name the five sodium salts of chlorine.

a. HClO
b. HClO₂
c. HClO₃
d. HClO₄
e. NH₄⁺

**Answer a**

+1

**Answer b**

+3
Exercise \(\PageIndex{9}\)

Name the five sodium salts of chlorine.

a. \(\text{SO}_3\)
b. \(\text{SO}_3^{\text{-2}}\)
c. \(\text{H}_2\text{SO}_3\)
d. \(\text{SO}_4^{\text{-2}}\)
e. \(\text{H}_2\text{SO}_4\)

Answer a

+6

Answer b

+4

Answer c

+4

Answer d

+6

Answer e

+6

Contributors

Robert E. Belford (University of Arkansas Little Rock; Department of Chemistry). The breadth, depth and veracity of this
work is the responsibility of Robert E. Belford, rebelford@ualr.edu. You should contact him if you have any concerns. This material has both original contributions, and content built upon prior contributions of the LibreTexts Community and other resources, including but not limited to:

- November Palmer & Ronia Kattoum (UALR)
- anonymous