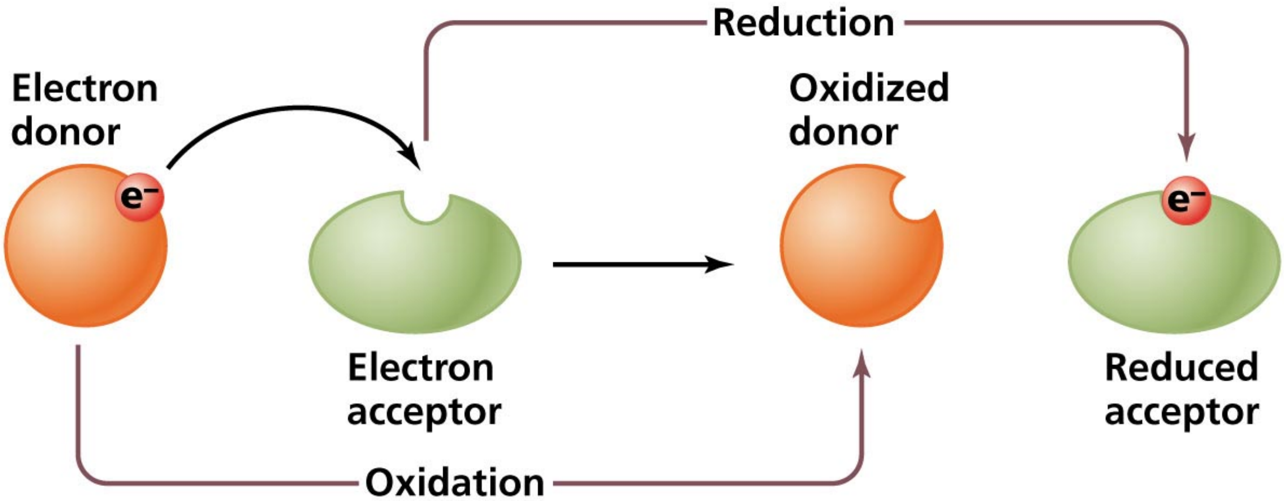


Unit 7 - Redox Reactions and Electrochemistry

7.1 - Introduction to Redox Reactions

pages 635-642 in Matter and Change

7.1 - Introduction to Redox Reactions

Redox is short for 'oxidation and reduction,' which are two complimentary types of chemical reactions.

- **Oxidation** originally referred to a reaction where substances combine with oxygen. These types of reactions are commonly called **combustion** or **corrosion**. For example, burning a log or rusting iron.
- **Reduction** originally referred to converting metal ores into pure metals by reducing the mass of the ore.

These two terms have broader meanings today.

The key to all redox reactions is that there is an **exchange of electrons.** That is, when one substance loses electrons, another one gains them.

For example,

$Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$

Since we have aqueous substances, the ions are dissociated... An ionic equation shows us
that:

$Cu_{(s)} + 2Ag^+_{(aq)} + 2NO_3^-{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + 2NO_3^-{}_{(aq)} + 2Ag_{(s)}$

Now we can remove the spectator ions and observe the net ionic equation.

$Cu_{(s)} + 2Ag^+_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + 2Ag_{(s)}$

7.1 - Introduction to Redox Reactions

$Cu_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$

What do we see now?

1. Copper began as a neutral atom, but became an ion with a +2 charge. An atom becomes positive by **losing** electrons:

   $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^{-}$

   Here, we say that copper has been **oxidized** because it has **lost electrons** (the electrons appear on the product side of the equation).

2. Silver was converted from an ion with a +1 charge to a neutral atom. Therefore, it must have gained an electron:

   $Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)}$

   We say silver has **gained** electrons and has been **reduced** (electrons appear on the reactant side of the equation).

Now, looking back at the net ionic equation, we can see the electrons gained by silver came from copper.

Note that a substance can **not** lose electrons if there is not another substance to take them. That is **oxidation cannot occur without reduction**.

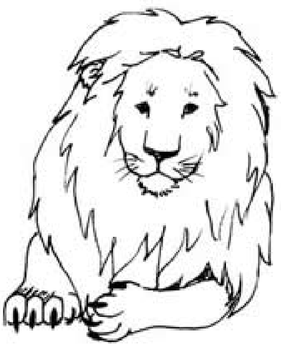
This exchange of electrons is what defines an **oxidation-reduction reaction (redox).**

Here is a mnemonic:

LEO

the Lion
Says

GER



LEO

Loss of
Electrons
is Oxidation

GER

Gain of
Electrons
is Reduction

Oxidation Numbers

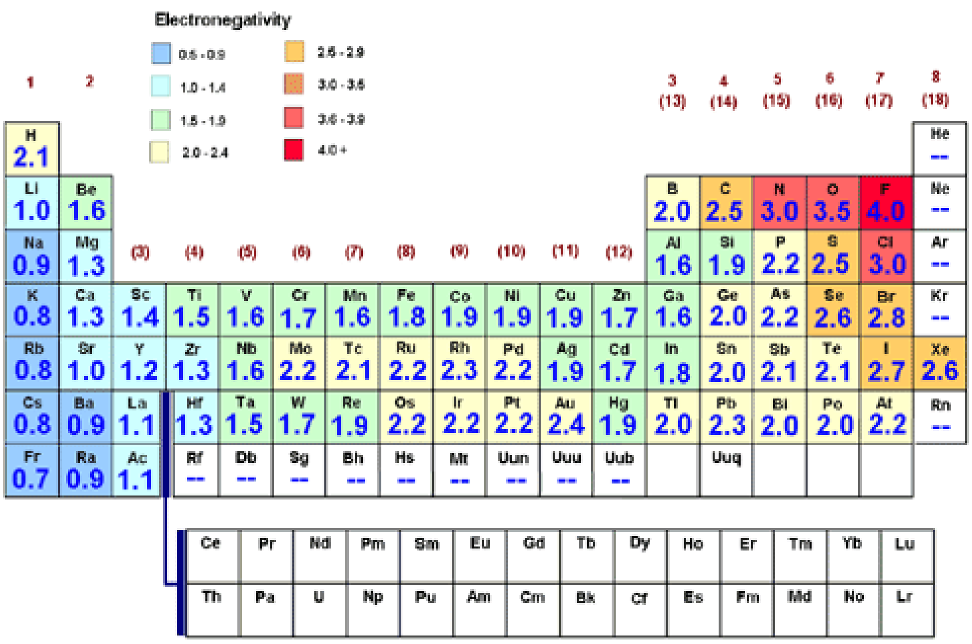
Since redox reactions involve exchanges of electrons between substances, there should be an ordered system to keep track of what gains, what loses,and how many electrons are involved. This record keeping system is called **oxidation numbers**.

Oxidation numbers are easiest to understand through ionic compounds because we already know there is an exchange of electrons.

However, electronegativity can help us understand oxidation numbers in covalent compounds. REMINDER: **Electronegativity** is the ability of a bonded atom to attract electrons towards itself. *(Table 6))*

-metals have a relatively low electronegativity
-non-metals have a high electronegativity

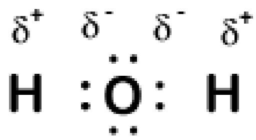
When two atoms are bonded together by a covalent bond, there is a tug ofwar for the electrons they are sharing due to differences in each atom's electronegativity. The more electronegative atom will pull harder



Consider the water molecule. Oxygen has a higher electronegativity than hydrogen. Therefore, the oxygen pulls the shared electrons closer to itself.

This close proximity of oxygen and the electrons makes oxygen partially negative.

Likewise, the part of the hydrogen that has lost the electron has become partially positive, while the side closest to oxygen has become partially negative.



We use the δ^+ to show the parts of the molecule that are partially positive and the δ^- to show partially negative parts.

We can think of each H 'losing' an electron and the O as 'gaining' two electrons (even though they are technically SHARING the electrons).

Charges given to atoms in this way are called **oxidation numbers**.

For water, hydrogen has an oxidation number of +1 because each has 'lost' an electron. Oxygen has an oxidation number of -2 because it has 'gained' two electrons.

Ex) What are the oxidation numbers for H_2O_2?

Helpful Hint: There are **common oxidation numbers for elements listed in *Table 17*

7.1 - Introduction to Redox Reactions

We will not need to draw Lewis structures often for redox reactions. However, knowing oxidation numbers will be key to our understanding of how these reactions work.

Use the **oxidation number rules** sheet when dealing with redox reactions *(Table 19)*.

- oxidation number of element in stable state = 0
- sum of oxidation numbers for polyatomic compound is 0 unless otherwise specified with accompanying charge

Note that oxidation numbers refer to **each atom** in a molecule or compound, and not the **total** atoms for a given element.

Ex) What is the oxidation number of each element in:

a) $SnCl_4$

b) HI

c) $Mg(NO_2)_2$

d) MgH_2

An increase in the oxidation number of an atom/ion represents oxidation.

A decrease in the oxidation number of an atom/ion represents reduction.

ex) Is Nitrogen oxidized or reduced when NH_3 --> N? How do you know?

7.1 Introduction to Redox Reactions Assignment

		Oxidation Numbers for each Element			
a.	$SnCl_4$	Sn ______	Cl ______		
b.	Ca_3P_2	Ca ______	P ______		
c.	SnO	Sn ______	O ______		
d.	Ag_2S	Ag ______	S ______		
e.	HI	H ______	I ______		
f.	N_2H_4	N ______	H ______		
g.	Al_2O_3	Al ______	O ______		
h.	S_8	S ______			
i.	HNO_2	H ______	N ______	O ______	
j.	O_2	O ______			
k.	H_3O^+	H ______	O ______		
l.	ClO_3^-	Cl ______	O ______		
m.	$S_2O_3^{2-}$	S ______	O ______		
n.	$KMnO_4$	K ______	Mn ______	O ______	
o.	$(NH_4)_2SO_4$	N ______	H ______	S ______	O ______

2. Determine the oxidation number of carbon in each of the following compounds:

   a. methane, CH_4

   b. formaldehyde, CH_2O

   c. carbon monoxide, CO

   d. carbon dioxide, CO_2

3. When elemental iron is made from Fe_2O_3, it iron oxidized or reduced?

4. Determine which of the following processes are oxidations and which are reductions:

a. Co^{2+} becomes Co

b. $2I^-$ becomes I_2.

c. Fe^{3+} becomes Fe^{2+}

d. Sn^{2+} becomes Sn^{4+}

5. Determine if each of the following changes is an oxidation, a reduction, or neither:

a. $SO_3^{2-} \rightarrow SO_4^{2-}$

b. $CaO \rightarrow Ca$

c. $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$

d. $2I^- \rightarrow I_2$

e. $IO_3^{1-} \rightarrow I_2$