Chemistry Lecture #104: Oxidation-Reduction

Oxidation-reduction (aka redox) reactions: A chemical reaction where electrons are transferred from one atom to another.

$$Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$$

In the above reaction, 2 electrons (or $2e^-$) are transferred from the bromine ions to the chlorine molecule.

Bromine loses electrons → bromine has been oxidized

Chlorine gains electrons → chlorine has been reduced

Oxidizing agent: the substance that steals electrons.

Chlorine steals electrons from bromine,

-or-

Chlorine oxidizes bromine. Chlorine is the oxidizing agent.

Reducing agent: The substance that donates electrons.

Bromine donates electrons to chlorine

-or-

Bromine reduces chlorine. Bromine is the reducing agent.

Chlorine oxidizes bromine and becomes reduced.
Bromine reduces chlorine and becomes oxidized.

The oxidizing agent is the substance containing the element that is reduced. The reducing agent contains the element that is oxidized.

To identify oxidizing and reducing agents, we need to know the rules for finding the oxidation numbers of atoms in a compound.

Rules

1. The oxidation number of atoms in a free element is zero.
   E.g., the atoms in Zn, H_2, S_8, and O_2 all have oxidation values of zero.

2. Oxidation number of hydrogen is always +1.
   E.g., for H_2S, H has an oxidation number of +1.

3. Most of the time, the oxidation number of oxygen is -2.
   E.g., in H_2O, oxygen has an oxidation number of -2.
   But in peroxides like H_2O_2 and Na_2O_2, oxygen has an oxidation value of -1.

4. The sum of the oxidation numbers must equal the charge of that compound.

5. The oxidation number of Al is +3. Those of groups 1A are +1. Group 2A has a value of +2.

What is the oxidation number of the elements in Na_2SO_4?

Solution

Na is a group IA element, so Na = +1
Oxygen is -2, so O = -2.

Since the sum of all the oxidation numbers must equal the charge of the compound, we can use algebra to find the charge on S. Na_2SO_4 is a neutral compound, so the total charge is zero.

$$\begin{aligned} Na + O + S &= 0 \\ 2(+1) + 4(-2) + S &= 0 \\ 2 + (-8) + S &= 0 \\ S &= +6 \end{aligned}$$

What is the oxidation number of the elements in $Cr_2O_7^{2-}$?

Solution

Oxygen has a value of -2, and the sum of all the charges in the compound is also -2. We use algebra to find the charge on Cr.

$$\begin{aligned} Cr + O &= -2 \\ 2Cr + 7(-2) &= -2 \\ 2Cr - 14 &= -2 \\ 2Cr &= 12 \\ Cr &= +6 \end{aligned}$$

The change in oxidation number of atoms allows us to identify oxidizing and reducing agents in a reaction.

Identify the oxidizing and reducing agents in the reaction

$Fe_2O_3 \quad + \quad 3CO \quad \rightarrow \quad 2Fe \quad + \quad 3CO_2$

Solution

Ignore the coefficients in front of each substance. We only want the oxidation numbers of the products and reactants. I've written the oxidation numbers of the atoms above the substance.

$$\overset{+3\ -2}{Fe_2O_3} \quad + \quad 3\ \overset{+2\ -2}{CO} \quad \rightarrow \quad 2\ \overset{0}{Fe} \quad + \quad 3\ \overset{+4\ -2}{CO_2}$$

Notice that the iron in Fe_2O_3 changes its oxidation state from +3 to 0 when it becomes elemental Fe.

$$\overset{+3\ -2}{Fe_2O_3} \rightarrow \overset{0}{Fe}$$

Thus, the iron atom has gone from Fe^{+3} to Fe. It has stolen 3 electrons. Fe_2O_3 is the oxidizing agent.

Notice that the carbon in CO changes its oxidation state from +2 to +4.

$$\overset{+2\ -2}{CO} \rightarrow \overset{+4\ -2}{CO_2}$$

Thus, the carbon atom has gone from C^{+2} to C^{+4}. Atoms gain positive charge when they lose electrons. Thus, CO is the reducing agent. It has donated electrons to the iron atom.