Oxidation-Reduction Reactions

Oxidation-reduction reactions or redox reactions are reactions in which electrons are transferred from one atom or molecule to another. There are many types of redox reactions the most common are single replacement and combustion reactions.

In this type of single replacement reaction it is relatively easy to see where the electrons are going.

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

Reactions that involve the loss of e^- are oxidation.

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

Reactions that involve the gaining of e^- are reduction.

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

Oxidation involves loss, reduction involves gain. OIL RIG

Combustion Reactions

In many redox reactions, such as combustion, it can be much more difficult to see how the electrons are moving around.

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + \ 2\ H_2O_{(g)}$$

What we need is a system to keep track of the number of electrons each atom has gained or lost. To do this we use oxidation numbers.

Oxidation Numbers

To keep track of the electrons in more complex reactions we use a system called oxidation numbers. We use a set of rules to assign oxidation numbers or states to each of the atoms in the reaction. We can then identify which atoms have gained or lost electrons according to the changes in oxidation numbers.

There are 6 rules for assigning oxidation states:

1- The oxidation state of an atom in it's elemental state is 0

2- The oxidation state of a monoatomic ion is the same as its charge.

3- In compounds, fluorine is always assigned an oxidation state of -1.

4- Oxygen is usually assigned an oxidation state of -2, exceptions include peroxides where oxygen has an oxidation state of -1.

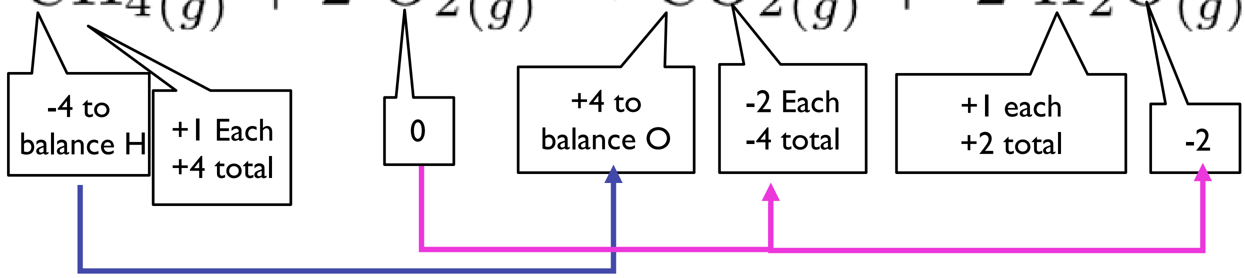
5- Hydrogen is usually assigned an oxidation state of+1, exceptions include metal hydrides where hydrogen has a -1 oxidation state.

6- The sum of the oxidation states must equal the charge on the compound.

Redox Reactions

Now we can look at a more complex redox reactions.

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + \ 2\ H_2O_{(g)}$$



Now we can identify the changes that have taken place.

$$CH_{4(g)} \rightarrow CO_{2(g)} \ + 8\ e^-$$

The carbon in the methane has lost 8 e^-, this is the oxidation reaction and methane is the reducing agent.

$$2\ O_{2(g)} + 8\ e^- \rightarrow CO_{2(g)} + \ 2\ H_2O_{(g)}$$

The oxygen has gained 8 electrons this is the reduction reaction and the oxygen is the oxidizing agent.

Ex:

Balancing Redox Reactions

$$\underset{+7}{\mathrm{Mn}}\underset{-2}{\mathrm{O}}_{4\,(aq)}^{-} + \underset{+2}{\mathrm{Fe}}^{2+}_{(aq)} \rightarrow \underset{+3}{\mathrm{Fe}}^{3+}_{(aq)} + \underset{+2}{\mathrm{Mn}}^{2+}_{(aq)}$$

1- Identify and write equations for each of the half-reactions.

$$\mathrm{MnO_4}^{-}_{(aq)} \rightarrow \mathrm{Mn}^{2+}_{(aq)} \qquad \mathrm{Fe}^{2+}_{(aq)} \rightarrow \mathrm{Fe}^{3+}_{(aq)}$$

2a- Balance all elements except for H and O

2b- Balance O using water.

$$\mathrm{MnO_4}^{-}_{(aq)} \rightarrow \mathrm{Mn}^{2+}_{(aq)} + 4\ \mathrm{H_2O}_{(l)} \qquad \mathrm{Fe}^{2+}_{(aq)} \rightarrow \mathrm{Fe}^{3+}_{(aq)}$$

2c- Balance H using H^+

$$8\ \mathrm{H}^{+}_{(aq)} + \mathrm{MnO_4}^{-}_{(aq)} \rightarrow \mathrm{Mn}^{2+}_{(aq)} + 4\ \mathrm{H_2O}_{(l)} \qquad \mathrm{Fe}^{2+}_{(aq)} \rightarrow \mathrm{Fe}^{3+}_{(aq)}$$

2d- Balance charge with e^-

$$5\ e^- + 8\ \mathrm{H}^{+}_{(aq)} + \mathrm{MnO_4}^{-}_{(aq)} \rightarrow \mathrm{Mn}^{2+}_{(aq)} + 4\ \mathrm{H_2O}_{(l)} \qquad \mathrm{Fe}^{2+}_{(aq)} \rightarrow \mathrm{Fe}^{3+}_{(aq)} + e^-$$

3- Balance e^- gained and lost

x5

4- Add half reactions, and cancel appropriately.

$$5\ e^- + 8\ \mathrm{H}^{+}_{(aq)} + \mathrm{MnO_4}^{-}_{(aq)} + 5\ \mathrm{Fe}^{2+}_{(aq)} \rightarrow 5\ \mathrm{Fe}^{3+}_{(aq)} + \mathrm{Mn}^{2+}_{(aq)} + 4\ \mathrm{H_2O}_{(l)} + 5\ e^-$$

5- Check that elements and charge balance.

Ex:

Future Thoughts- What other phases could reactions take place in?