Redox Reactions

In Chemistry there are type main reaction types, acid-base reactions, which we will cover later in the year, and redox reactions. Redox is short for **Red**uction-**Ox**idation reactions. Reduction and Oxidation are like lending and borrowing, you cannot have one without the other. Redox reactions are all around us and they are occurring at all times. Examples of redox reactions include:

- The we eat to maintain our body
- Growth of plants
- Batteries
- The burning of fuels
- Extraction and corrosion of metals
- Photography
- Bushfires
- Explosions

The definition of what is a redox reaction is has changed over the years and has therefore included many more reactions. The first definition of redox was

Oxidation is a gain of oxygen
Reduction is a loss of oxygen

The burning of magnesium is oxidation, as you can see the magnesium gains oxygen.

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$$

The removal of oxygen from iron ore is a reduction process, however as you can see carbon monoxide also gains oxygen in the process and therefore has undergone oxidation

$$Fe_2O_{3(s)} + 3CO_{(s)} \longrightarrow 2Fe_{(l)} + 3CO_{2(g)}$$

However as they continued to research different reactions they saw that that this definition had become too limited as there were other reactions that were considered to be redox that did not include oxygen in the reaction at all. Let us look at the example below. Even though there is oxygen here in the equation and the carbon has oxidised, it has also lost hydrogen.

$$CH_{4(g)} + 2O_2 \longrightarrow CO_{2(g)} + 2H_2O_{(g)}$$

This can be further illustrated by the following equation. Here no oxygen is present, however there is a gain of hydrogen. This was thought to be a redox reaction also and hence the definition of what redox is had to be extended.

$$C_2H_{4(g)} + H_{2(g)} \longrightarrow C_2H_{6(g)}$$

Oxidation is a gain of oxygen or loss of hydrogen
Reduction is a loss of oxygen or a gain of hydrogen

Finding this fact opens up a whole new lot of reactions that are considered to be redox. However if it can be extended in this way maybe it can be extended further. To do this we look at the mechanism of the reaction by looking at the electron transfers in these reactions to see if this is linked to our previous definitions. Let us take the reaction below.

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$$

In this case we have an ionic compound form as magnesium is a metal and oxygen is a non-metal. Therefore the electrons are swapped from one atom to another. Since magnesium is a metal it is going to be the one that loses electrons to the oxygen. This can be broken down into the reactions below.

$$Mg - 2e^- \longrightarrow Mg^{2+}$$

$$O + 2e^- \longrightarrow O^{2-}$$

These reactions above are known as **half reactions** as they are the two steps or mini reactions that occur to make up the whole reaction. The half reaction for magnesium remains the same regardless of which of the non-metals it reacts with. Since we know from above that the magnesium is oxidised in this reaction we can come to the conclusion that substances that are oxidised will in fact lose electrons. This leads to the next definition of redox which is a more broad and will work regardless whether or not there is oxygen and/or hydrogen in the reaction.

Oxidation is a loss of electrons Reduction is a gain of electrons

If you have trouble remembering this property of redox reactions you can remember the following

O	xidation
I	s
L	oss
R	eduction
I	s
G	ain

Take the following reaction between magnesium and a dilute acid

$$2H^+_{(aq)} + Mg_{(s)} \longrightarrow Mg^{2+}_{(aq)} + H_{2(g)}$$

In this case we have can see that the hydrogen has gone from a positive ion to neutral and therefore must have gained electrons, therefore it has undergone reduction. On the other hand the magnesium has gone from a neutral charge to a positive one and therefore must have lost electrons, hence oxidation has occurred.

The two half reactions are as follows

$$2H^+_{(aq)} + 2e^- \longrightarrow H_{2(g)}$$

$$Mg_{(s)} - 2e^- \longrightarrow Mg^{2+}_{(aq)}$$

This definition of redox in terms of electron transfer does again have its limitations, it can only be used for ionic compounds as they are the only ones that transfer electrons. However there are many reactions that are redox that are ionic and therefore we again must come up with a better definition of what constitutes a redox reaction.

We know in covalent molecules due to differences in electronegativity there is an unequal sharing of electrons in molecules making them polar. This polarity can also be defined as a partial transfer of electrons, which would indicate a definition much like the electron transfer one above. However figuring out the total effect of all of the polarity changes can be complicated so a way of keeping track of these was devised. This system is known as oxidation numbers. Once oxidation numbers have been assigned it is easy to see if redox has occurred. If there is a change in the oxidation for a particular atom between the reactants and the products then redox has occurred.

There are a set of rules that must be followed when assigning oxidation numbers.

1. **The sum of all the oxidation numbers of the atoms in a molecule is equal to zero and in an ion it is equal to the charge on the ion.**

   Eg *For H_2O it is a molecule with no charge therefore the sum of the oxidation numbers of oxygen and hydrogen must equal zero*

   Eg *For CO_3^{2-} it is a polyatomic ion and therefore the sum of the oxidation numbers of carbon and oxygen must be equal to the charge, which is –2.*

2. **Certain atoms always have the same oxidation number in their compounds**
   - Group 1 metals oxidation number = +1
   - Group 2 metals oxidation number = +2
   - Oxygen has an oxidation number = -2 (except in peroxides = -1)
   - Hydrogen has an oxidation number = +1 (except with metals = -1)

For all other elements other than those mentioned the oxidation numbers can be figured out using the rules mentioned above

We can figure out oxidation numbers in the following way. Lets take nitric acid HNO_3 as an example. We want to find out the oxidation number of the nitrogen. There is no overall charge on the molecule so the total oxidation number must be zero

0	=	Oxidation No. (H)	+	Oxidation No. (N)	+	3 x Oxidation No. (O)
0	=	+1	+	Oxidation No. (N)	+	3 x –2
0	=	1	+	Oxidation No. (N)	–	6
0	=	-5	+	Oxidation No. (N)		
0 + 5	=	Oxidation No. (N)				
5	**=**	**Oxidation No. (N)**				

Therefore we can now extend our definition of redox

Oxidation is an increase in oxidation number
Reduction is a decrease in oxidation number

But how can we can we see if a redox reaction has occurred or not. Lets take the following equation and assign oxidation numbers which are shown above each atom.

$$\overset{0}{Cu_{(s)}} + \overset{+1}{2Ag^{+}_{(aq)}} \longrightarrow \overset{+2}{Cu^{2+}_{(aq)}} + \overset{0}{2Ag_{(s)}}$$

We can see in this case that the oxidation numbers for each element have changed and therefore a redox reaction has occurred. We can also see that the copper has increased in oxidation number and has therefore oxidised whilst the silver has undergone reduction as the oxidation number decreased

For the reaction below we can see that it is not a redox reaction as there is no change in oxidation numbers

$$\overset{+1-1}{HCl_{(g)}} + \overset{+1-2}{H_2O} \longrightarrow \overset{+1-2}{H_3O^{+}} + \overset{-1}{Cl^{-}_{(aq)}}$$

Balancing half reactions and redox equations

Half reactions as shown previously are not always as easy and simple as they seem above. At times they are quite complicated to be able to write and balance the half reactions correctly. As always with any reaction two things must balance the number of each type of atom and the charge.

Lets take the following half reaction for a redox reaction

$$Cr_2O_7^{2-} \longrightarrow Cr^{3+}$$

How we balance this equation is not immediately obvious. We can balance the Cr atoms relatively easily, however the O atoms are only on one side of the equation and not on the other so we can't balance them like we have learnt to so far. Again we use a set of rules that we apply to be able to balance both the atoms and the charge lets balance the equation above.

1. Balance all other atoms other than oxygen and hydrogen

$$Cr_2O_7^{2-} \longrightarrow \mathbf{2}Cr^{3+}$$

As you can see we balance the Cr atoms by adding a coefficient of 2 on the Right hand side (RHS)

2. Balance the oxygen atoms by adding an equivalent number of water molecules

$$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + \mathbf{7H_2O}$$

We have balanced the 7 O atoms on the LHS by adding 7 water molecules to the RHS

3. Balance the hydrogen atoms by adding an equivalent number of H^+ ions.

$$Cr_2O_7^{2-} + \mathbf{14H^{+}} \longrightarrow 2Cr^{3+} + 7H_2O$$

We have balanced the 14 H atoms in the water on the RHS by adding 14 H^+ ions to the LHS

4. Balance the charge by adding electrons (e-)

$$Cr_2O_7^{2-} + 14H^{+} + \mathbf{6e^{-}} \longrightarrow 2Cr^{3+} + 7H_2O$$

Since total charge on LHS = +12 and on RHS = + 6 we need to equal it out by adding 6 electrons to LHS

Once we have balanced the two half reactions we then need to combine them and balance the overall redox reaction. Since it is a redox reaction one reaction must be reduction and the other oxidation. Therefore one half reaction must have electrons on the LHS and the other must have them on the RHS. The electrons in the half equations must balance, you cannot have one reaction giving 5 electons if the other one can only accept 2. Lets take the example of the half equations shown below

$$MnO_4^-{}_{(aq)} \quad + \quad 8H^+{}_{(aq)} \quad + \quad 5e^- \longrightarrow Mn^{2+}{}_{(aq)} \quad + \quad 4H_2O$$

$$Fe^{2+}{}_{(aq)} \longrightarrow Fe^{3+}{}_{(aq)} \quad + \quad e^-$$

As we can see here the top half reaction has 5 electons whilst the bottom one has 1 electron. We even this up by multiplying the whole bottom equation by 5

$$MnO_4^-{}_{(aq)} \quad + \quad 8H^+{}_{(aq)} \quad + \quad 5e^- \longrightarrow Mn^{2+}{}_{(aq)} \quad + \quad 4H_2O$$

$$\mathbf{5}Fe^{2+}{}_{(aq)} \longrightarrow \mathbf{5}Fe^{3+}{}_{(aq)} \quad + \quad \mathbf{5}e^-$$

We can now combine the two reactions

$$MnO_4^-{}_{(aq)} \quad + \quad 8H^+{}_{(aq)} \quad + \quad 5e^- \longrightarrow Mn^{2+}{}_{(aq)} \quad + \quad 4H_2O$$

$$5Fe^{2+}{}_{(aq)} \longrightarrow 5Fe^{3+}{}_{(aq)} \quad + \quad 5e^-$$

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

Now we cancel out all the things that are common to both sides of the equation to give our final equation

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

Final reaction equation is

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

By checking the atoms we can see that they balance on both sides and by checking the charge we can see it is +17 on both sides, therefore this redox reaction has been balanced correctly

Review Questions

Identifying oxidation or reduction	Pg 82	Q 12, 13
Oxidation numbers	Pg 85 – 87	Q 14 – 19
Identifying redox reactions	Pg 88 – 89	Q 22
Half Reactions	Pg 87 – 88	Q 20, 21
Reactions	Pg 89	Q 23