Lesson 4 – Oxidation numbers

An Oxidation number is the arbitrary assignment of a number to an element that reflects its gain or loss of electrons. Assignment of oxidation numbers to elements is an easier way to see if an element is oxidised or reduced without having to write half equations especially if it involves non-metal elements.

For example – $NO_3^- \rightarrow NO_2$ looking at this reaction it is difficult to know whether nitrogen is oxidised or reduced. On the other hand $Fe \rightarrow Fe^{3+}$ is a lot easier to identify as oxidation as we can see electrons have been lost.

The concept of an atom's **oxidation number** or **oxidation state** is based on the following set of rules:

1. **An atom of a free element** has an oxidation number of zero.

   For example, Cl in Cl_2 has an oxidation state of 0, as does Na atoms in Na(s) and F atoms in F_2.

2. **A monatomic ion** has an oxidation number equal to its charge.

   For example, Na^+ has an oxidation number of +1, while Al^{3+} has an oxidation number or oxidation state of +3 and N^{3-} will have an oxidation number or state of -3.

3. **Hydrogen** has an oxidation number of +1,

   The only exception is when it forms a compound with a metal (metal hydride) such as LiH where its oxidation state is -1.

4. **Group 1 metals** have an oxidation state of +1 while **group 2 metals** have an oxidation state of +2

5. **Oxygen** has an oxidation number of -2

   The exception is in peroxides such as H_2O_2 where oxygen has an oxidation number of -1

   and in F_2O where it has an oxidation state of +1.

6. **The sum of the oxidation numbers** for all atoms in a polyatomic compound is equal to the charge on the compound. For example:

   - MnO_4^- - the oxidation number of Mn + 4 X the oxidation number of O = -1

   - H_2CO_3 - 2 X the oxidation number of hydrogen + the oxidation number of C + 3 X oxidation number of O = 0

7. The oxidation number of fluorine is always –1. Chlorine, bromine, and iodine usually have an oxidation number of –1, unless they're in combination with an oxygen or fluorine.

**Note* - *oxidation numbers are written with the sign before the number eg -2 or +1 . This is different for when we write the charges on an ion such as* Cl^{1-} *or* Mn^{5+}.

1. Workout the oxidation number of the element that is underlined below.
   a. $\underline{Cl}O_3^-$
   Cl + 3 X -2 = -1 => Cl has an oxidation state of +5
   b. $\underline{Mn}O_2$
   Mn + 2 X -2 = 0 => Mn has an oxidation state of +4
   c. $\underline{S}O_4^{2-}$
   S + 4 X -2 = -2 => S has an oxidation state of +6
   d. $K\underline{Mn}O_4$
   1+ Mn + 4 X -2 = 0 => Mn has an oxidation state of +7
   e. $\underline{S}O_3$
   S + 3 X -2 = 0 => S has an oxidation state of +6
   f. $\underline{Cr}(H_2O)_6^{3+}$
   Cr + 12 X 1 + 6 X -2 = +3 => Cr has an oxidation state of +3
   g. $\underline{Cr}Cl_3$
   Cr + 3 X -1 = 0 => Cr has an oxidation state of +3
   h. $\underline{Mn}O_4^-$
   Mn + 4 X -2 = -1 => Mn has an oxidation state of +7
   i. $\underline{Al}$
   Al = 0
   j. $\underline{Mo}O_4^{2-}$
   Mo + 4 X -2 = -2 => Mo has an oxidation state of +6
   k. $\underline{Cl}_2$
   Cl = 0
   l. $Na\underline{Cl}O_4$

   1 + Cl + 4 X -2 = 0 => Cl has an oxidation state of +7

   m. $Na\underline{Cl}O_2$
   1 + Cl + 2 X -2 = 0 => Cl has an oxidation state of +3
   n. $\underline{Ce}_2O_4^-$
   2 X Ce + 4 X -2 = -1 => Ce has an oxidation state of +3.5
   o. $\underline{S}_8$
   S= 0
   p. $\underline{O}_3$
   O = 0
   q. $\underline{Al}$(s) Aluminium metal
   Al = 0
   r. $\underline{I}O_4^-$
   I + 4 X -2 = 0 => I has an oxidation state of +7
   s. $\underline{I}_2$
   I = 0
   t. $\underline{I}F_7$
   I + 7 X -1 = 0 => I has an oxidation state of +7
   u. $H\underline{I}O_4$
   1 + I + 4 X -2 = 0 => I has an oxidation state of +7

2. Oxidation, involves an increase in oxidation number while reduction involves a decrease in oxidation number.
In each of the changes below identify if the reaction represents oxidation or reduction .
The first one is done for you.
$Cl_2(g) \rightarrow 2Cl^-(s)$
Cl goes from an oxidation state of 0 to -1. Cl is therefore reduced.

a. $MnO_4^- \rightarrow Mn^{2+}$

Mn goes from an oxidation state of +7 to +2. Mn is therefore reduced.

b. $SO_4^{2-} \rightarrow SO_2$

S goes from an oxidation state of +6 to +4. S is therefore reduced.

c. $ClO_3^- \rightarrow Cl_2$

Cl goes from an oxidation state of +5 to 0. Cl is therefore reduced.

d. $Al(s) \rightarrow Al^{3+}(aq)$

Al goes from an oxidation state of 0 to +3. Al is therefore oxidised.

e. $Cr^{3+} \rightarrow Cr_2O_7^{2-}$

Cr goes from an oxidation state of +3 to +6. Cr is therefore oxidised.

3. Consider the following reactions. Indicate if the reaction is a redox reaction and if so identify the oxidant and reductant by reference to a change in oxidation state. The first one is done for you.
$3Ag^+(aq) + Al(s) \rightarrow Al^{3+} + 3Ag(s)$
This is a redox reaction.
oxidant – $Ag^+(aq)$ oxidation state +1 changes to Ag (s) with an oxidation state of 0. It is reduced.
Reductant – Al(s) oxidation state of 0 change to $Al^{3+}(aq)$ with an oxidation state of +3. It is oxidised.

a. $Zn(s) + 2Au^+(aq) \rightarrow Au(s) + Zn^{2+}(aq)$
This is a redox reaction.
oxidant – $Au^+(aq)$ oxidation state +1 changes to Au (s) with an oxidation state of 0. It is reduced.
Reductant – Zn(s) oxidation state of 0 change to $Zn^{2+}(aq)$ with an oxidation state of +2. It is oxidised.

b. $C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
This is a redox reaction.
oxidant – C_2H_6 (aq) oxidation state of carbon is +3 and changes to +4 in CO_2. It is oxidised.
Reductant – $O_2(g)$ oxidation state of 0 changes to -2 in CO_2 and H_2O . It is reduced.

c. $2Co^{3+}(aq) + Ni(s) \rightarrow 2Co^{2+}(aq) + Ni^{2+}(aq)$

This is a redox reaction.
oxidant – $Co^{3+}(aq)$ oxidation state +3 changes to $CO^{2+}(aq)$ with an oxidation state of +2. It is reduced.
Reductant – Ni(s) oxidation state of 0 change to $Zn^{2+}(aq)$ with an oxidation state of +2. It is oxidised.

d. $Cu(s) + 2NO_3^-(aq) + 4H^+(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$

This is a redox reaction.
oxidant – $NO_3^-(aq)$ – nitrogen has an oxidation state of +5 changes to +4 in $NO_2(g)$. It is reduced.
Reductant – Cu(s) oxidation state of 0 change to $Cu^{2+}(aq)$ with an oxidation state of +2. It is oxidised.

e. $2MnO_4^-(aq) + 6I^-(aq) + 4H_2O(l) \rightarrow 2MnO_2(s) + 3I_2(aq) + 8OH^-(aq)$

This is a redox reaction.
oxidant – $MnO_4^-(aq)$ - Mn has an oxidation state of +7 changes to +4 MnO_2. It is reduced.
Reductant – $I^-(aq)$ - oxidation state of -1 changes to 0 in $I_2(aq)$. It is oxidised.

f. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

This is not a redox reaction.
No change in oxidation numbers takes place.
Ag^+ remains Ag^+ in AgCl(s)
Cl^- remains Cl^- in AgCl(s)

g. $6CO_2(g) + 6H_2O(l) \rightarrow C_2H_{12}O_6(aq) + 6O_2(g)$

This is a redox reaction.
Reductant – $H_2O(l)$ - O has an oxidation state of -2 changes to 0 in O_2. It is oxidised.
Oxidant – $CO_2(g)$ - oxidation state of C in CO_2 changes to +4 to 0 in $C_6H_{12}O_6(aq)$. It is reduced.

4. Identify the conjugate pairs in the redox reactions below.

reductant *oxidant*

a. $Zn(s) + 2Au^+(aq) \rightarrow Au(s) + Zn^{2+}(aq)$

oxidant *reductant*

$Zn(s)/Zn^{2+}(aq)$ $Au^+(aq)/Au(s)$

reductant *oxidant*

b. $C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

oxidant *reductant*

$C_2H_6(g)$ / $CO_2(g)$ *$O_2(g)$ / $H_2O(l)$*

oxidant *reductant*

c. $2Co^{3+}(aq) + Ni(s) \rightarrow 2Co^{2+}(aq) + Ni^{2+}(aq)$

reductant *oxidant*

$Co^{3+}(aq)$ / $Co^{2+}(aq)$ *$Ni(s)$ / $Ni^{2+}(aq)$*

reductant *oxidant*

d. $Cu(s) + 2NO_3^-(aq) + 4H^+(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$

oxidant *reductant*

$Cu(s)$ / $Cu^{2+}(aq)$ *$NO_3^-(aq)$ / $NO_2(g)$*

oxidant *reductant*

e. $2MnO_4^-(aq) + 6I^-(aq) + 4H_2O(l) \rightarrow 2MnO_2(s) + 3I_2(aq) + 8OH^-(aq)$

reductant *oxidant*

$MnO_4^-(aq)/MnO_2(s)$ *$I^-(aq)/I_2(aq)$*

oxidant *reductant*

f. $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + CO_2(g)$

reductant *oxidant*

$Fe_2O_3(s)/Fe(s)$ *$CO(g)/CO_2(g)$*

oxidant *reductant*

g. $CuO(s) + Mg(s) \rightarrow MgO(s) + Cu(s)$ *$CuO(s)$ /$Cu(s)$* *$Mg(s)$ / $MgO(s)$*

reductant *oxidant*