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. $\underline{Cl}O_3^-$
    b. $\underline{Mn}O_2$
    c. $\underline{S}O_4^{2-}$
    d. $K\underline{Mn}O_4$
    e. $\underline{S}O_3$
    f. $\underline{Cr}(H_2O)_6^{3+}$
    g. $\underline{Cr}Cl_3$
    h. $\underline{Mn}O_4^-$
    i. $\underline{Al}$
    j. $\underline{Mo}O_4^{2-}$
    k. $\underline{Cl}_2$
    l. $Na\underline{Cl}O_4$
    m. $\underline{V}_2O_3$
    n. $Na\underline{Cl}O_2$
    o. $\underline{Ce}_2O_4^-$
    p. $\underline{S}_8$
    q. $\underline{O}_3$
    r. $\underline{Al}(s)$ Aluminium metal
    s. $\underline{I}O_4^-$
    t. $\underline{I}_2$
    u. $\underline{I}F_7$
    v. $H\underline{I}O_4$

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

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

   c. $ClO_3^- \rightarrow Cl_2$

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

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

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 of particular elements. 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. $Zn(s) + 2Au^+(aq) \rightarrow 2Au(s) + Zn^{2+}(aq)$

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

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

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

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

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

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

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

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

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

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

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

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

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

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

h. $CuO(s) + Mg(s) \rightarrow MgO(s) + Cu(s)$