Chem12 Oxidation and Reduction Reactions : W.S. - 10

The words **oxidize** and **oxidization** used to mean "to react with oxygen". For example, the rusting of iron which is the combining of iron with oxygen is an oxidation reaction.

e.g. $Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$ (rust)

But if iron combines with sulfur we also have an oxidation reaction. The modern definition of oxidation is this : A species is **oxidized** if it loses electrons. This can only happen in the presence of an **oxidizing agent** which is the species that takes the electrons.

The words **reduce** and **reduction** used to refer to reducing a metal ore. This means removing the oxygen atoms from the compound to obtain pure metal.

e.g. $Fe_2O_3(s) + C(s) \rightarrow Fe(l) + CO_2(g)$ + energy

In the above reaction the $Fe_2O_3(s)$ is the iron ore. It is ground up and mixed with coal (carbon) at a high temperature and the reaction proceeds with liquid iron being collected at the bottom of the mixture. The modern definition of reduction is this : A species is **reduced** if it gains electrons. This can only happen in the presence of a **reducing agent** which is the species that gives up electrons.

Oxidation cannot occur without reduction. Chemical reactions of this type are called **oxidation-reduction reactions** or **redox reactions**.

In the above example, Fe is reduced and C is oxidized.

e.g. When a piece of copper wire is placed in a solution of silver nitrate, $Ag(NO_3)$ (aq), the solution begins to turn blue as $Cu^{2+}(aq)$ forms , and shiny crystals of Ag(s) form on the copper. The nitrate ions are spectator ions and don't contribute the overall reaction.

$Ag^+(aq) + Cu(s) \rightarrow Ag(s) + Cu^{2+}(aq)$ (unbalanced)

Cu(s) is being oxidized (it is the reducing agent) and Ag^+ is being reduced. (it is the oxidizing agent)

This redox reaction is made up of two **half-reactions** .

$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$	reduction	equation #1
$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-}$	oxidation	equation #2

The complete reaction is the sum of 2 x equation #1 plus equation #2. The number of electrons on the reactant side must equal the number of electrons on the product side. The electrons lost by a species in a redox reaction must equal the electrons gained by another species. Both atoms and charge must be balanced in a redox reaction. The balanced reaction is given below.

$$2Ag^{+}(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq) \quad \text{(balanced)}$$

Oxidation Numbers

In redox reactions, **oxidation numbers** are assigned to each atom in a species. The reason for this will be seen later.

•The oxidation number for an atom in its elemental form is zero.

•Alkali elements (H, Li, Na....) always have an oxidation number of +1 (except for H in metallic hydrides like NaH. In this case the oxidation number is -1).

•Alkaline Earths (Ca, Mg, ...) always have an oxidation number of +2.

•The oxidation number of oxygen is usually -2. (except for peroxides such as H_2O_2 where it is -1).

•The oxidation number of the halogens (F, Cl,...) is usually -1.

•The total oxidation number of a species equals the charge on that species and is the sum of the oxidation numbers of the component atoms.

•Ionic compounds should be separated into their component ions when assigning oxidation numbers.

Some atoms can have several different oxidation numbers. Oxidation numbers are usually (but not always) integers. **<u>Important note</u>** : **In a redox reaction, the oxidation number of the species oxidized increases and the oxidation number of the species reduced is decreased**.

Example : Find the oxidation numbers for nitrogen in these species.

a) NH_4^+ Total O.N. = +1. O.N. for H = +1 (for each H atom) therefore the O.N. for N = -3.

b) NO_3^- Total O.N. = -1. O.N. for O = -2 (for each O atom) therefore the O.N. for N = +5

Example : Find the oxidation number for boron.

a) $H_4B_3O_7^-$ Total O.N. = -1. O.N. for H = +1 (for each H atom) O.N. for O = -2 for each O atom. Therefore the O.N. for B = +3. [4(H) + 3(B) + 7(O) = -1]

b) B Total O.N. = 0. O.N. for B = 0. The O.N. for an atom in its elemental form is always zero.

Problems :

1) Find the oxidation numbers for all atoms in the following species :

a) MnO_2	b) SO_3^{2-}	c) N_2O_3	d) Cl_2
e) OH^-	f) Fe^{3+}	g) ClO_3^-	h) CaH_2
i) $AlBO_3$	j) $Cr_2O_7^{2-}$	k) H_2O_2	l) Ca

2) In the equation below, has iron been oxidized or reduced? Has carbon been oxidized or reduced? Give the oxidizing and reducing agents.

$Fe_2O_3 + C \rightarrow Fe + CO_2$ (unbalanced)

3) In the equation below, has iodine been oxidized or reduced ? Has chlorine been oxidized or reduced ? (hint; If a species is oxidized, its oxidation number increases)

$I_2 + 5Cl_2 + 6H_2O \rightarrow 2HIO_3 + 10HCl$

4) In the reactions below (not necessarily balanced), state whether each is a half-reaction, redox reaction or neither. State the species being oxidized and reduced.

a) $Zn + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2$

b) $Sn^{2+} + 2e^- \rightarrow Sn$

c) $Zn + NO_3^- \rightarrow Zn^{2+} + NO$

d) $6H^+ + 6e^- + ClO_3^- \rightarrow Cl^- + 3H_2O$

e) $CO + H_2O \rightarrow CO_2 + H_2$

f) $NH_4HCO_3 \rightarrow NH_3 + H_2O + CO_2$

g) $Mn^{2+} + 4H_2O \rightarrow MnO_4^- + 8H^+ + 5e^-$

h) $2KClO_3 \rightarrow 2KCl + 3O_2$

Answers : 1)a) Mn, +4, O, -2, b) S, +4, O, -2, c) N, +3, O, -2, d) Cl, 0, e) O, -2, H, +1, f) Fe, +3, g) Cl, +5, O, -2, h) Ca, +2, H, -1, i) Al, +3, B, +3, O, -2, j) Cr, +6, O, -2, k) H,+1, O, -1, l) Ca, 0, 2) Fe has been reduced from +3 to 0. C has been oxidized from 0 to +4. C is the reducing agent, Fe is the oxidizing agent, 3) I has been oxidized, O.N. increases from 0 to +5, Cl has been reduced, O.N. decreases from 0 to -1, 4 a) redox, Zn is oxidized, H is reduced, b) half reaction, Sn is reduced, c) redox, Zn is oxidized, N is reduced, d) half-reaction, Cl is reduced, e) redox, C is oxidized, H is reduced, f) neither, oxidation numbers don't change, g) half-reaction, Mn is oxidized, h) redox, Cl is reduced, O is oxidized.