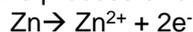


## 1.7 Redox

Oxidation is the process of **electron loss**:



*It involves an increase in oxidation number*

Reduction is the process of **electron gain**:



*It involves a decrease in oxidation number*

### Rules for assigning oxidation numbers

- All uncombined elements have an oxidation number of zero
- The oxidation numbers of the elements in a compound add up to zero
- The oxidation number of a monoatomic ion is equal to the ionic charge
- In a polyatomic ion ( $\text{CO}_3^{2-}$ ) the sum of the individual oxidation numbers of the elements adds up to the charge on the ion
- Several elements have invariable oxidation numbers in their common compounds.

eg. Zn,  $\text{Cl}_2$ ,  $\text{O}_2$ , Ar all have oxidation numbers of zero

In NaCl Na = +1 Cl = -1  
Sum = +1 -1 = 0

e.g.  $\text{Zn}^{2+}$  = +2 Cl<sup>-</sup> = -1

e.g. in  $\text{CO}_3^{2-}$  C = +4 and O = -2  
sum = +4 + (3 x -2) = -2

Group 1 metals = +1

Group 2 metals = +2

Al = +3

H = +1 (except in metal hydrides where it is -1 eg NaH)

F = -1

Cl, Br, I = -1 except in compounds with oxygen and fluorine

O = -2 except in peroxides ( $\text{H}_2\text{O}_2$ ) where it is -1 and in compounds with fluorine.

We use these rules to identify the oxidation numbers of elements that have variable oxidation numbers.

What is the oxidation number of Fe in  $\text{FeCl}_3$

Using rule 5, Cl has an O.N. of -1

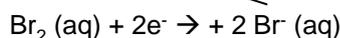
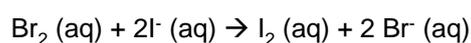
Using rule 2, the O.N. of the elements must add up to 0

Fe must have an O.N. of +3

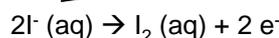
in order to cancel out 3 x -1 = -3 of the Cl's

Note the oxidation number of Cl in  $\text{CaCl}_2$  = -1 and not -2 because there are two Cl's  
Always work out the oxidation per one atom of the element.

### Redox equations and half equations



Br has reduced as it has gained electrons



I has oxidised as it has lost electrons

*A reduction half equation only shows the parts of a chemical equation involved in reduction. The electrons are on the left*

*An oxidation half equation only shows the parts of a chemical equation involved in oxidation. The electrons are on the right*

The oxidising agent is Bromine water. It is an **electron acceptor**

The reducing agent is the iodide ion. It is an **electron donor**

**reducing agents are electron donors**

**oxidising agents are electron acceptors**

*An oxidising agent (or oxidant) is the species that causes another element to oxidise. It is itself reduced in the reaction.*

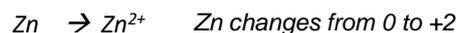
*A reducing agent (or reductant) is the species that causes another element reduce. It is itself oxidised in the reaction..*

When naming oxidising and reducing agents always refer to full name of substance and not just name of element or ion

## Balancing redox equations

### Writing half equations

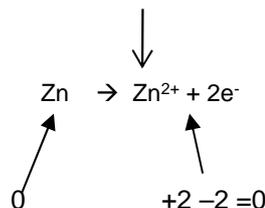
1. Work out oxidation numbers for element being oxidised/ reduced



2. Add electrons equal to the change in oxidation number

For reduction add e's to reactants

For oxidation add e's to products



3. check to see that the sum of the charges on the reactant side equals the sum of the charges on the product side

### More complex half-equations

If the substance that is being oxidised or reduced contains a varying amount of O (eg  $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$ ) then the half equations are balanced by adding  $\text{H}^+$ ,  $\text{OH}^-$  ions and  $\text{H}_2\text{O}$ .

In acidic conditions use  $\text{H}^+$  and  $\text{H}_2\text{O}$

**Example:** Write the half equation for the change  $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$

1. Balance the change in O.N. with electrons

*Mn changes from +7 to +2*

*Add 5 electrons to reactants*



2. Add  $\text{H}_2\text{O}$  in products to balance O's in  $\text{MnO}_4^-$



3. Add  $\text{H}^+$  in reactants to balance H's in  $\text{H}_2\text{O}$



4. check to see that the sum of the charges on the reactant side equals the sum of the charges on the product side

$$-1 + 8 - 5 = +2$$

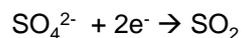
$$+2$$

**Example:** Write the half equation for the change  $\text{SO}_4^{2-} \rightarrow \text{SO}_2$

1. Balance the change in O.N. with electrons

*S changes from +6 to +4*

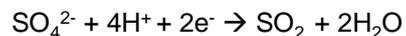
*Add 2 electrons to reactants*



2. Add  $\text{H}_2\text{O}$  in products to balance O's in  $\text{SO}_4^{2-}$



3. Add  $\text{H}^+$  in reactants to balance H's in  $\text{H}_2\text{O}$



4. check to see that the sum of the charges on the reactant side equals the sum of the charges on the product side

$$-4 + 4 = 0$$

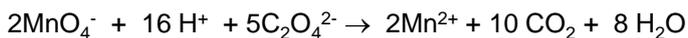
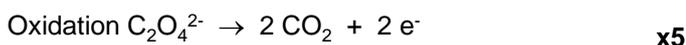
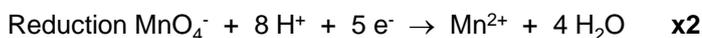
$$0$$

### Combining half-equations

To make a full redox equation combine a reduction half equation with a oxidation half equation

To combine two half equations there must be equal numbers of electrons in the two half equations so that the electrons cancel out

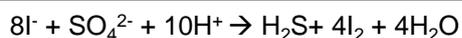
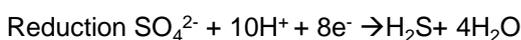
#### Example 1



Multiply the half equations to get equal electrons

Add half equations together and cancel electrons

#### Example 2



Multiply the half equations to get equal electrons

Add half equations together and cancel electrons