2. Redox

oxidation is the process **of electron loss**: $Zn \rightarrow Zn^{2+} + 2e^-$ *It involves an increase in oxidation number*	reduction is the process **of electron gain**: $Cl_2 + 2e^- \rightarrow 2Cl^-$ *It involves a decrease in oxidation number*

Rules for assigning oxidation numbers

1. All uncombined elements have an oxidation number of zero

eg . Zn, Cl_2, O_2, Ar all have oxidation numbers of zero

2. The oxidation numbers of the elements in a compound add up to zero

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

3. The oxidation number of a monoatomic ion is equal to the ionic charge

e.g. Zn^{2+} = +2 Cl^- = -1

4. In a polyatomic ion (CO_3^{2-}) the sum of the individual oxidation numbers of the elements adds up to the charge on the ion

e.g. in CO_3^{2-} C = +4 and O = -2
sum = +4 + (3 x -2) = -2

5. Several elements have invariable oxidation numbers in their common compounds.

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 (H_2O_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 $FeCl_3$

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 $CaCl_2$ = -1 and not -2 because there are two Cl's
Always work out the oxidation for one atom of the element

Naming using oxidation number

If an element can have various oxidation numbers then the oxidation number of that element in a compound can be given by writing the number in roman numerals

$FeCl_2$: Iron (II) chloride
$FeCl_3$ Iron (III) chloride
MnO_2 Manganese (IV) Oxide

In IUPAC convention the various forms of sulfur,nitrogen and chlorine compounds where oxygen is combined are all called sulfates, nitrates and chlorates with relevant oxidation number given in roman numerals. If asked to name these compounds remember to add the oxidation number.

NaClO: sodium chlorate(I)
$NaClO_3$: sodium chlorate(V)
K_2SO_4 potassium sulfate(VI)
K_2SO_3 potassium sulfate(IV)

$NaNO_3$ sodium nitrate (V)
$NaNO_2$ sodium nitrate (III)

Redox equations and half equations

$$Br_2 (aq) + 2I^- (aq) \rightarrow I_2 (aq) + 2 Br^- (aq)$$

$Br_2 (aq) + 2e^- \rightarrow + 2 Br^- (aq)$

Br has reduced as it has gained electrons

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

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

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

$2I^- (aq) \rightarrow I_2 (aq) + 2 e^-$

I has oxidised as it has lost electrons

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

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

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

reducing agents are electron donors

oxidising agents are electron acceptors

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

Redox Reactions

metals generally form ions by losing electrons with an increase in oxidation number to form positive ions:
$Zn \rightarrow Zn^{2+} + 2e^-$

non-metals generally react by gaining electrons with a decrease in oxidation number to form negative ions
$Cl_2 + 2e^- \rightarrow 2Cl^-$

Oxygen is reducing because its oxidation number is decreasing from 0 to -2

$$\overset{0}{4Li} + \overset{0}{O_2} \rightarrow 2\overset{+1}{Li_2}\overset{-2}{O}$$

Lithium is oxidising because its oxidation number is increasing from 0 to +1

Tungsten is reducing because its oxidation number is decreasing from +6 to 0

$$\overset{+6}{WO_3} + \overset{0}{3H_2} \rightarrow \overset{0}{W} + 3\overset{+1}{H_2}O$$

Hydrogen is oxidising because its oxidation number is increasing from 0 to +1

Nitrogen is reducing because its oxidation number is decreasing from +5 to+4

$$2Sr(\overset{+5}{N}\underset{-2}{O_3})_2 \rightarrow 2SrO + 4\overset{+4}{NO_2} + \overset{0}{O_2}$$

Oxygen is oxidising because its oxidation number is increasing from -2 to 0

Note that not all the oxygen atoms are changing oxidation number in this reaction

Chlorine is reducing because its oxidation number is decreasing from +1 to -1

$$2\ \overset{-3}{N}H_3 + NaCl\overset{+1}{O} \rightarrow \overset{-2}{N_2}H_4 + NaC\overset{-1}{l} + H_2O$$

Nitrogen is oxidising because its oxidation number is increasing from -3 to -2

Redox Reactions of Metals and acid

ACID + METAL → SALT + HYDROGEN

Hydrogen is reducing because its oxidation number is decreasing from +1 to 0

+1 0

$2HCl + Mg \rightarrow MgCl_2 + H_2$

0 +2

Magnesium is oxidising because its oxidation number is increasing from 0 to +2

Be able to write equations for reactions of metals with hydrochloric acid and sulphuric acid

$Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$

Observations: These reaction will effervesce because H_2 gas is evolved and the metal will dissolve

Disproportionation

Disproportionation is the name for a reaction where an element in a single species simultaneously oxidises and reduces.

$Cl_2(aq) + H_2O(l) \rightarrow HClO(aq) + HCl(aq)$

Chlorine is both simultaneously reducing and oxidising changing its oxidation number from 0 to -1 and 0 to +1

$2Cu^+ \rightarrow Cu + Cu^{2+}$

Copper(I) ions (+1) when reacting with sulphuric acid will disproportionate to Cu^{2+} (+2) and Cu (0) metal

Balancing Redox equations

Writing half equations

1. Work out oxidation numbers for element being oxidised/ reduced — *$Zn \rightarrow Zn^{2+}$* *Zn changes from 0 to +2*

2. Add electrons equal to the change in oxidation number
   For reduction add e's to reactants
   For oxidation add e's to products

   $Zn \rightarrow Zn^{2+} + 2e^-$

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

   0   +2 –2 =0

More complex Half equations

If the substance that is being oxidised or reduced contains a varying amount of O (eg $MnO_4^- \rightarrow Mn^{2+}$) then the half equations are balanced by adding H^+, OH^- ions and H_2O.

In acidic conditions use H^+ and H_2O

Example: Write the half equation for the change $MnO_4^- \rightarrow Mn^{2+}$

1. Balance the change in O.N. with electrons — *Mn changes from +7 to +2*, *Add 5 electrons to reactants* — $MnO_4^- + 5e^- \rightarrow Mn^{2+}$
2. Add H_2O in products to balance O's in MnO_4^- — $MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$
3. Add H^+ in reactants to balance H's in H_2O — $\boxed{MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O}$
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 $SO_4^{2-} \rightarrow SO_2$

1. Balance the change in O.N. with electrons — *S changes from +6 to +4*, *Add 2 electrons to reactants* — $SO_4^{2-} + 2e^- \rightarrow SO_2$
2. Add H_2O in products to balance O's in SO_4^{2-} — $SO_4^{2-} + 2e^- \rightarrow SO_2 + 2H_2O$
3. Add H^+ in reactants to balance H's in H_2O — $\boxed{SO_4^{2-} + 4H^+ + 2e^- \rightarrow SO_2 + 2H_2O}$
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

Reduction $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ **x2**

Oxidation $C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$ **x5**

Multiply the half equations to get equal electrons

$\boxed{2MnO_4^- + 16H^+ + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O}$

Add half equations together and cancel electrons

Example 2

Reduction $SO_4^{2-} + 10H^+ + 8e^- \rightarrow H_2S + 4H_2O$

Oxidation $2I^- \rightarrow I_2 + 2e^-$ **x4**

Multiply the half equations to get equal electrons

$\boxed{8I^- + SO_4^{2-} + 10H^+ \rightarrow H_2S + 4I_2 + 4H_2O}$

Add half equations together and cancel electrons