Oxidation numbers

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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 (eg Fe^{2+}) is equal to the ionic charge

Eg 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

Eg in CO_3^{2-} C = +4 and O = -2

sum = +4 + (3 x -2) = -2

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 the underlined element in the following compound ?

$\underline{Fe}Cl_3$

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

The compound is Iron(III) chloride

What is the oxidation number of the underlined element in the following compound ?

$\underline{Mn}O_2$

Using rule 5, O has an O.N. of –2

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

Mn must have an O.N. of +4
in order to cancel out 2 x –2 = -4 of the O's

The compound is Manganese (IV) oxide

What is the oxidation number of the underlined element in the following ion ?

<u>Mn</u>O_4^-

Using rule 5, O has an O.N. of –2

Using rule 4, the O.N. of the elements must add up to -1

Mn must have an O.N. of +7
in order to cancel out 4 x –2 = -8 of the O's
+7 -8 = -1 (the charge on the ion)

The ion is the Manganate (VII) ion

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

What is the oxidation number of the underlined element in the following compounds and ions ?

Compound/ ion	Oxidation number
$\underline{C}O_2$	+4
$\underline{N}H_3$	-3
$\underline{Mg}Cl_2$	+2
$\underline{N}H_4^+$	-3
$\underline{S}O_4^{2-}$	+6
$\underline{Cr}O_4^{2-}$	+6
$\underline{Cr}_2O_7^{2-}$	+6

Definitions

Oxidation takes place when an element in a reaction

- Increases its oxidation number
- Loses electrons

Reduction takes place when an element in a reaction

- Decreases its oxidation number
- Gains electrons

To help you remember this, you can use "OIL RIG" – **O**xidation **is** **L**oss **R**eduction **is** **G**ain (of electrons)

Redox Reactions

In fact, oxidation never takes place on its own - nor does reduction.

When one substance is oxidised in a reaction, another one is reduced.

A **Redox** reaction is one in which both reduction and oxidation take place.

To work out which element is oxidised and which is reduced in a reaction, we go through these steps:

1 Write down the chemical equation for the reaction.
2 Go through and work out the oxidation numbers for every element in the equation.
3 Look for an element that has **increased** its oxidation number from one side of the equation to the other - it has been **oxidised**.
4 Look for an element that has **decreased** its oxidation number from one side of the equation to the other - it has been **reduced**.

I has oxidised,
increased in O.N

$$\underset{0}{Br_2}\,(aq) + 2\underset{-1}{I^-}\,(aq) \rightarrow \underset{0}{I_2}\,(aq) + 2\,\underset{-1}{Br^-}\,(aq)$$

Br has reduced,
decreased in O.N

Definition:

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

Oxidising agents **accept electrons**.

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**.

Note that the oxidising or reducing agent is the **compound** involved in the reaction - not one of the elements in it.

AQA want reducing and oxidising agents defined in terms of electron transfer

I has oxidised,
increased in O.N

$$\underset{0}{Br_2}\,(aq) + 2\underset{-1}{I^-}\,(aq) \rightarrow \underset{0}{I_2}\,(aq) + 2\,\underset{-1}{Br^-}\,(aq)$$

Br has reduced,
decreased in O.N

The oxidising agent is Bromine water (refer to full name of substance and not just name of element)
The reducing agent is the Iodide ion.

Half equations

Half equations can be written to show just the oxidation or reduction, and include electrons.

oxidation half equation:

$2I^- \rightarrow I_2 + 2e^-$ (electrons are lost)

reduction half equation:

$Br_2 + 2e^- \rightarrow 2Br^-$ (electrons are gained)

Exam Questions

1. Deduce the oxidation numbers of iodine in IO_3^-, I^- and I_2.

2. The following is an equation for a redox reaction.

$2NO + 12H^+ + 10I^- \rightarrow 2NH_4^+ + 2H_2O + 5I_2$

(i) Define *oxidation* in terms of electrons.

(ii) Deduce the oxidation state of nitrogen in NO and of nitrogen in NH_4^+

(iii) Identify the species formed by oxidation in this reaction

Exam Questions 2

(a) In terms of electron transfer, what does the reducing agent do in a redox reaction?

(b) What is the oxidation state of an atom in an uncombined element?

(c) Deduce the oxidation state of nitrogen in each of the following compounds.

(i) NCl_3

(ii) Mg_3N_2

(iii) NH_2OH