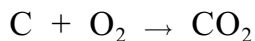


TOPIC 11.

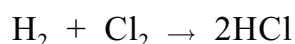
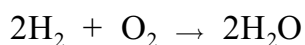
OXIDATION AND REDUCTION REACTIONS.

Early concepts of oxidation.

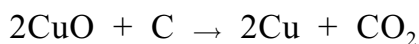
The name "oxidation" was initially applied to reactions where substances combined with the element oxygen. Thus any substance burning in air was said to be oxidised, the product being some type of oxide. For example, burning carbon to produce carbon dioxide is an oxidation, as shown by the equation



Subsequently it was realised that reactions of substances with elements other than oxygen were of essentially the same type. For example, hydrogen can react with oxygen to form the compound water, but equally it can react with chlorine to form the compound hydrogen chloride. In both reactions the free element hydrogen is converted to a compound of hydrogen and another non-metal, and so both were classed as oxidations even though no oxygen was involved in the second case.



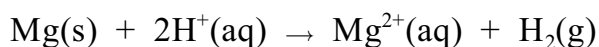
The reverse reaction, conversion of compounds such as oxides of metals to the elemental metal were called "reduction" reactions, for example, the reduction of copper(II) oxide to copper by heating with charcoal (carbon).



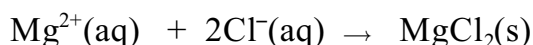
The gain or loss of oxygen is still a useful way of recognising some oxidation or reduction reactions, but with a knowledge of the structure of atoms, a rather different definition is now more widely used.

Oxidation reactions as a loss of electrons.

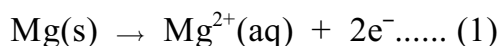
Consider the following reaction in which the metal, magnesium, is treated with hydrochloric acid, as discussed in Topic 6. The magnesium dissolves to form Mg^{2+} ions in solution and hydrogen gas is evolved. The equation given previously was



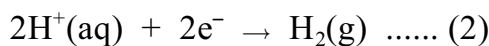
Notice that the free element magnesium (Mg) has been converted to the compound magnesium chloride, MgCl_2 , which is present in solution as its component ions $\text{Mg}^{2+}(\text{aq})$ and $\text{Cl}^-(\text{aq})$. Evaporating the water would isolate this compound as shown in the following equation.



This is therefore just like the examples of oxidation given above where elements were converted to compounds. However, when the reaction equation is written in this way the electronic nature of the change is more apparent. The electrically neutral Mg atoms are converted to the charged Mg^{2+} cations. For this to occur, each Mg atom has lost 2 electrons according to the following **HALF EQUATION**:



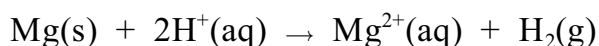
At the same time, the H^+ ions have gained 1 electron per ion to form hydrogen atoms which have become H_2 molecules according to the following half equation.



In this second half reaction the element hydrogen (H_2) has been formed.

It can be seen from this example that an oxidation reaction is one in which a species loses electrons and a reduction reaction is one in which a species gains electrons. As electrons lost by one species must be accepted by another, then both reactions must occur simultaneously. Thus equation (1) is an **OXIDATION HALF REACTION** (electrons lost by Mg) and equation (2) is a **REDUCTION HALF REACTION** (electrons gained by H^+).

The overall reaction equation is obtained by combining the separate half equations in such a way that the number of electrons lost exactly equals the number of electrons gained, and the final equation is called a **REDOX EQUATION**. In this example, the balanced overall equation results by simply adding the two half equations.



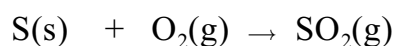
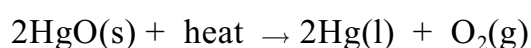
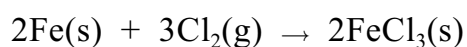
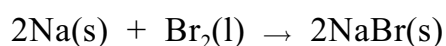
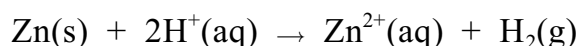
Note that the 2 electrons which appeared previously on each side of the half equations have now cancelled out.

All reactions of the type previously regarded as oxidations on the basis of gain of oxygen atoms have the same characteristic as in the magnesium example above - the transfer of electrons from the species oxidised to the species reduced. Thus the most general definitions of oxidation and reduction reactions are:

OXIDATION IS THE LOSS OF ELECTRONS
REDUCTION IS THE GAIN OF ELECTRONS

The species which causes the oxidation to occur (H^+ in the above example) is called the **OXIDIZING AGENT** or the **OXIDANT**, while the species which is oxidized is called the **REDUCING AGENT** or **REDUCTANT**. Equations written in the manner of (1) and (2) above are called **ION-ELECTRON HALF EQUATIONS**.

The following are all examples of redox reactions.



In each example, free elements were converted to compounds or alternatively, elements combined in compounds were converted back to the free state.

Check your understanding of this section.

Why is the combination of magnesium with either oxygen or chlorine regarded as the same type of reaction?

Define “oxidation” and “reduction” in terms of electron transfer.

Define the terms oxidant, reductant, oxidizing agent, reducing agent, ion-electron half equation.

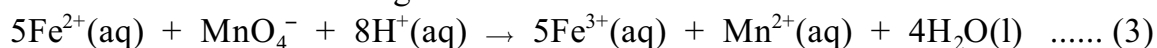
List some reduction reactions that are vital to our industrialised world.

Why must an oxidation reaction be accompanied by a reduction reaction?

What factor determines the stoichiometric ratio of reductant and oxidant in a balanced redox equation?

Could the conversion of one compound into another compound be a redox reaction? This proposition is answered in the following section.

Now consider the following reaction:



No free element has either been formed from or converted to a compound but the Fe^{2+} ion has obviously lost 1 electron to produce the Fe^{3+} ion. From the above definition of oxidation, this means that Fe^{2+} has been oxidised. We therefore need a method of recognising redox reactions which includes this situation where there are no free elements on either side of the equation.

The concept of oxidation number.

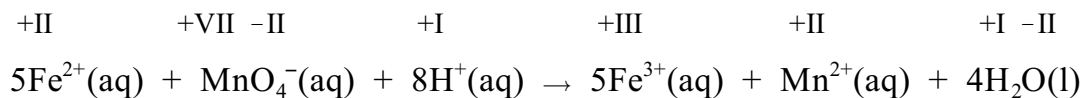
In the above example, the half equation for the reaction involving iron is



As electrons appear on the right hand side, this is clearly an oxidation. Since all oxidations are necessarily accompanied by a reduction, then which is the species reduced? The answer is not readily apparent because manganese is present on the left as part of the polyatomic ion MnO_4^- and on the right as the simple Mn^{2+} cation. Likewise, hydrogen is present on the left as the cation H^+ and on the right as part of the compound H_2O . There are no uncombined elements on either side to act as a guide. A clue is provided by noting that the Mn atom in the MnO_4^- ion has lost oxygen atoms in the reaction, leaving it as the Mn^{2+} ion - an indication of a reduction reaction. However, by assigning a quantity known as the **OXIDATION NUMBER** (or **OXIDATION STATE**) to each atom in the equation and observing which atoms change their oxidation numbers during the reaction, the species oxidised and reduced can be established unequivocally. Oxidation number is the **notional charge** that an atom would bear if all electrons in each bond to it were assigned to the more electronegative atom. Oxidation numbers are not ionic charges and, to avoid confusion, should always be written in Roman numerals with a + or - sign as a prefix. This also serves to distinguish oxidation numbers from valence states which have no sign.

The rules for assigning oxidation numbers are given in the table on page XI-5.

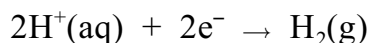
Using these rules we now return to the reaction (3) above and assign an oxidation number to each atom.



Starting with Fe^{2+} which we already know to be oxidised, note that its oxidation number has increased from +II to +III. Thus another definition of oxidation is:

A SPECIES IS OXIDISED IF ITS OXIDATION NUMBER INCREASES IN A REACTION.

Examination of all the other oxidation numbers shows that the only other species to change its oxidation number during the reaction is Mn from +VII in MnO_4^{-} to +II in Mn^{2+} . This is a decrease in oxidation number. As a species on the left must have been reduced, this suggests that when reduction occurs, the oxidation number of that species decreases. To check this, return to the half-reaction (2) which has already been designated as a reduction,

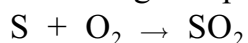


It can be seen that the oxidation number of H changes from +I to 0, i.e. a decrease occurs. This confirms that a reduction can be defined in terms of oxidation number change as:

A SPECIES IS REDUCED IF ITS OXIDATION NUMBER DECREASES IN A REACTION.

In reaction (3) above then, Mn in the MnO_4^{-} ion is the species reduced following its change in oxidation number from +VII to +II.

As another example, consider the reaction given previously -



The oxidation numbers of both S and O on the left are zero because they are elements. In the compound SO_2 , the oxidation number of O is -II (Rule 4). Then the oxidation number of the S atom in SO_2 must be +IV in order that the sum of the oxidation numbers in the compound adds up to zero (Rule 5). Therefore, S is the species oxidised (O.N. increases from zero to +IV) and O is the species reduced (O.N. decreases from zero to -II).

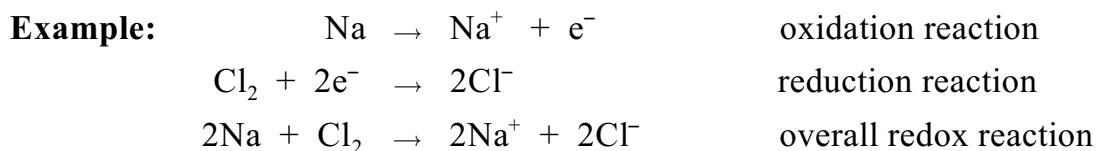
Rules for assigning oxidation numbers.

- The oxidation number of a simple cation is the same as the cationic charge.
e.g. Na^+ O.N. = +I; Mg^{2+} O.N. = +II; Al^{3+} O.N. = +III
Use was made of this in earlier Topics to name unambiguously cations such as those of copper which could be Cu^{2+} or Cu^+ and in one system of naming covalent compounds.
- The oxidation number of a simple anion is the same as the anionic charge.
e.g. O^{2-} O.N. = -II; Cl^- O.N. = -I; N^{3-} O.N. = -III
- The oxidation number of atoms in an element = zero.
- In all compounds except peroxides and OF_2 , oxygen has the oxidation number -II, and H in most compounds has oxidation number = +I.
- The sum of all the oxidation numbers in a polyatomic ion must equal the charge on the ion, or in a compound, must add up to zero.
e.g. in MnO_4^- , the sum of all oxidation numbers is $(4 \times -\text{II}) + Z = -1$ where Z is the oxidation number of Mn. Thus $Z = +\text{VII}$.
or, in the molecule CO_2 , $Y + (2 \times -\text{II}) = 0$, where Y stands for the oxidation number of carbon in the compound CO_2 . Thus $Y = +\text{IV}$.
Note that this does not imply that a C^{4+} ion is present in carbon dioxide - the carbon/oxygen bonds are in fact covalent.

Writing redox equations using ion-electron half equations.

The following procedure should be followed in writing redox equations using the ion-electron half equation method.

- Set out the skeleton half equations for the oxidation and the reduction.
- Balance the atoms on both sides of each half equation.
- Then, balance the charge on both sides of each half equation by **adding electrons as required**.
- Finally, combine the oxidation and reduction half equations in the ratio which will cancel out the electrons from each side.



Location of oxidants and reductants in the Periodic Table.

From the definition of oxidation as a loss of electrons, it is apparent that elements which have few electrons more than a noble gas in their outer levels will be the most easily oxidised - i.e. metals. Consequently, Group I elements are the easiest to oxidise as they have only 1 electron more than a noble gas structure and therefore the least amount of energy is required to remove that electron. Group II and Group III elements are also mostly easy to oxidise. These elements are therefore good reducing agents, as electrons can be removed from their outer levels with only a relatively small input of energy required. In contrast, elements in Group VII are very close to attaining the inert gas electron structure, requiring only the **gain** of 1 electron per atom. Thus the halogens are easily reduced and are therefore strong oxidizing agents as is also the element oxygen from Group VI where the gain of just 2 electrons allows the noble gas electron structure to be achieved.

Check your understanding of this section.

Define the term "oxidation number".

How is the oxidation number of an atom represented?

Define oxidation and reduction in terms of change in oxidation number.

How does oxidation number differ from valence of an atom?

Where in the Periodic Table are the strongest oxidants located?

Where in the Periodic Table are the strongest reductants located?

Why are metals usually oxidised while non-metals generally are reduced in redox reactions?

As a generalisation, the elements most easily oxidised are located on the left hand side of the Periodic Table while those most easily reduced (and therefore the strongest oxidants) are on the right hand side, excluding the noble gases.

Objectives of this Topic.

When you have completed this Topic, including the tutorial questions, you should have achieved the following goals:

1. Recognise oxidation reactions as:
 - (i) a free element converted to a compound, particularly by combination with oxygen and other non-metals.
 - (ii) a species losing electrons,
 - (iii) a species increasing its oxidation number.
2. Recognise reduction reactions in similar terms.
3. Be able to use ion electron half equations to balance simple redox equations.
4. Be able to deduce the oxidation number of any atom in a reaction.
5. Know the meaning of the terms oxidation; reduction; oxidant; oxidising agent; reductant; reducing agent.
6. Be aware of the location in the Periodic Table of the elements which are the easiest to oxidise and to reduce.

SUMMARY

Historically, early chemists observed that most elements combined with oxygen to form compounds and this type of reaction was called an oxidation. Later it was realised that other reactive non-metals such as the halogens enter into similar reactions but the same term was still applied. When the electronic nature of atoms was discovered, it was realised that the real basis for oxidation reactions was the transfer of electrons from the atom being oxidised to another atom, the latter being said to be reduced. As electrons must balance in reactions, it is fundamental that electrons lost by the species oxidised must be gained by the species being reduced and one cannot proceed without the other occurring simultaneously. The term redox reaction is used to reinforce this fact. Whenever an element is converted to a compound or vice-versa, the reaction is necessarily a redox reaction. The terms oxidising agent or oxidant are used to describe the species which causes another to be oxidised and is itself reduced in that process. Similarly, the terms reducing agent or reductant are applied to the species which is oxidised

However, many redox reactions only take place between compounds and no free elements are involved, the requirement that electrons be transferred being met without any free elements being consumed or produced. In order to allow such reactions to be easily identified as redox, an artificial method of bookkeeping of electrons in the atoms involved has been devised, called the oxidation number. This is a number which is assigned to each atom according to a set of rules. An atom undergoing an oxidation will experience an increase in its oxidation number while an atom being reduced will have its oxidation number decreased. Oxidation number is not the actual charge on an atom nor is it the valency of that atom. To distinguish oxidation numbers from charge or valency, Roman numerals should be used for all oxidation numbers.

Thus there are several methods of recognising a redox reaction:

- (i) as free elements and compounds being interconverted and especially where atoms of non-metals such as oxygen are gained or lost in that interconversion.
- (ii) as species gaining/losing electrons
- (iii) as species changing oxidation number.

Elements that are the most easily oxidised (good reducing agents) are the metals from the first few families as these require relatively little energy for the removal of their outer level electrons. Conversely, the elements that are the best oxidising agents come from the families that are only a few electrons short of the noble gas electron structure, particularly elements such as oxygen and the halogens.

TUTORIAL QUESTIONS - TOPIC 11.

1. Explain the following terms: oxidation half equation; reduction half equation; oxidizing agent; oxidant; reducing agent; reductant.

2. Give three definitions of "oxidation" and three definitions of "reduction". Illustrate each with an example.

3. Give the oxidation number of the underlined atom in each of the following:

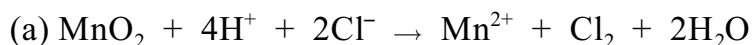
- | | | | |
|--|--|---|--|
| (i) $\underline{\text{K}}^+$ | (ii) $\underline{\text{Al}}$ | (iii) $\underline{\text{S}}\text{O}_4^{2-}$ | (iv) $\underline{\text{Fe}}\text{SO}_4$ |
| (v) $\underline{\text{N}}_2$ | (vi) $\underline{\text{Na}}\text{Cl}$ | (vii) $\underline{\text{Na}}_2\text{SO}_4$ | (viii) $\text{Na}_3\underline{\text{P}}\text{O}_4$ |
| (ix) $\text{Na}_2\underline{\text{C}}\text{O}_3$ | (x) $\underline{\text{Mg}}\text{O}$ | (xi) $\underline{\text{Mg}}\text{Cl}_2$ | (xii) $\underline{\text{N}}\text{H}_3$ |
| (xiii) $\underline{\text{N}}\text{O}_3^-$ | (xiv) $\underline{\text{N}}^{3-}$ | (xv) $\underline{\text{Cu}}\text{O}$ | (xvi) $\underline{\text{Cu}}_2\text{O}$ |
| (xvii) $\underline{\text{Fe}}_2\text{O}_3$ | (xviii) $\underline{\text{Cr}}_2\text{O}_7^{2-}$ | (xix) $\underline{\text{Bi}}^{3+}$ | (xx) $\underline{\text{C}}\text{O}_2$ |

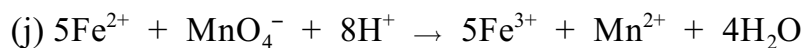
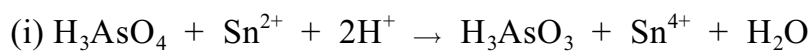
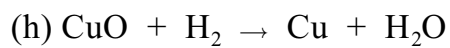
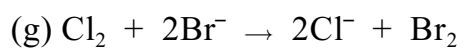
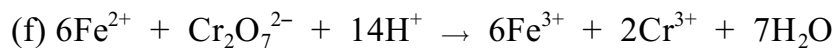
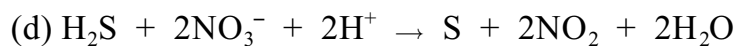
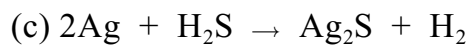
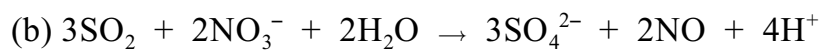
4. Using the ion electron half equation method, balance the following redox equations.

- (i) $\text{Fe}^{3+} + \text{I}^- \rightarrow \text{Fe}^{2+} + \text{I}_2$
- (ii) $\text{Sn}^{2+} + \text{Fe}^{3+} \rightarrow \text{Sn}^{4+} + \text{Fe}^{2+}$
- (iii) $\text{Cu} + \text{Ag}^+ \rightarrow \text{Cu}^{2+} + \text{Ag}$
- (iv) $\text{H}^+ + \text{Al} \rightarrow \text{H}_2 + \text{Al}^{3+}$
- (v) $\text{H}_2\text{S} + \text{Fe}^{3+} \rightarrow \text{H}^+ + \text{S} + \text{Fe}^{2+}$

5. In each of the following balanced redox equations, identify:

- (i) the atom oxidised and its product with their oxidation numbers
- (ii) the atom reduced and its product with their oxidation numbers.
- (iii) the reducing agent
- (iv) the oxidising agent

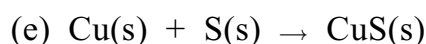
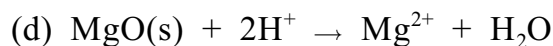
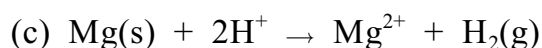
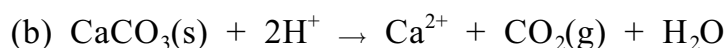
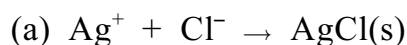




6. Give the oxidation number of the underlined atom in each of the following species:



7. Indicate which of the following equations represent redox reactions by writing "redox" after the equation. Where the reaction is redox, give the atom oxidised, the atom reduced and their oxidation numbers.



8. The mnemonic OILRIG is often quoted as a means of remembering what happens to electrons in oxidation and reduction reactions. What would this mnemonic be when expanded?

9. Sulfur dioxide reacts with water to form sulfite ion. Is this a redox reaction? Justify your answer.

10. Identify in which of the following species sulfur is in

(i) its lowest oxidation state

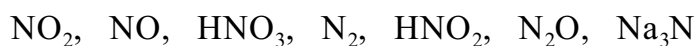
(ii) its highest oxidation state:



11. Identify in which of the following species nitrogen is in

(i) its lowest oxidation state

(ii) its highest oxidation state:



ANSWERS TO TUTORIAL TOPIC 11

1. Oxidation half equation: An equation showing only the oxidation part of a redox reaction, balanced with respect to both atoms and charge. The charge balance is achieved by adding the required number of electrons to the right hand side of the equation.

Reduction half equation: An equation showing only the reduction part of a redox reaction, balanced with respect to both atoms and charge. The charge balance is achieved by adding the required number of electrons to the left hand side of the equation.

Oxidizing agents are any species that **cause** oxidation and thus are themselves reduced - e.g. the oxidizing agent O_2 is reduced to O^{2-} when it oxidises a metal such as Zn to Zn^{2+} .

Oxidant: Another name for an oxidizing agent - oxidants cause an oxidation to occur and are themselves reduced in the redox reaction.

Reducing agents are any species that **cause** reduction and thus are themselves oxidized - e.g. the reducing agent H_2S is oxidized to S when it reduces Fe^{3+} to Fe^{2+} .

Reductant: Another name for a reducing agent - reductants cause a reduction to occur and are themselves oxidized in the redox reaction.

2. Oxidation of a species occurs when:

(i) it gains oxygen atoms e.g. $2Mg + O_2 \rightarrow 2MgO$

(ii) it loses electrons e.g. $Mg \rightarrow Mg^{2+} + 2e^-$

(iii) its oxidation number increases.

e.g. the oxidation number of Mg in both (i) and (ii) increases from zero on the LHS to +II on the RHS.

Reduction of a species occurs when:

(i) it loses oxygen atoms e.g. $CuO + H_2 \rightarrow Cu + H_2O$

(ii) it gains electrons e.g. $Cu^{2+} + 2e^- \rightarrow Cu$

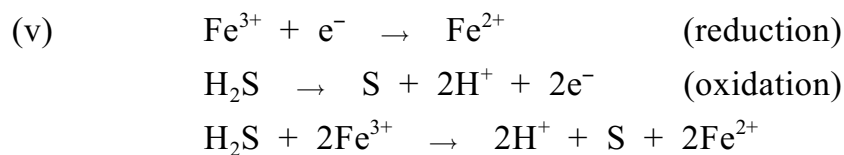
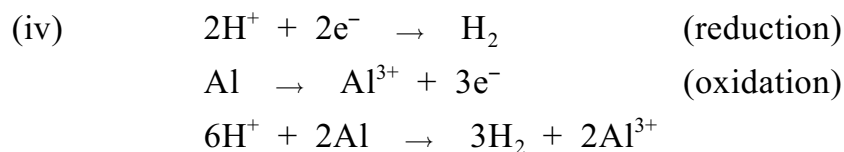
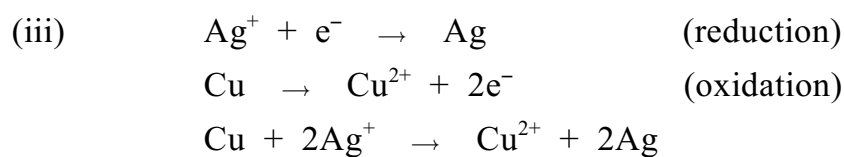
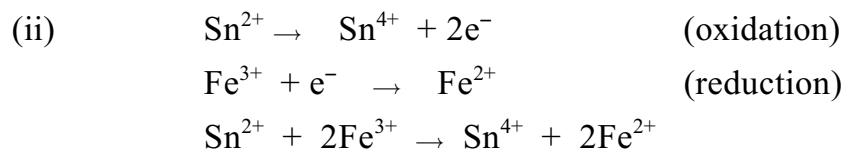
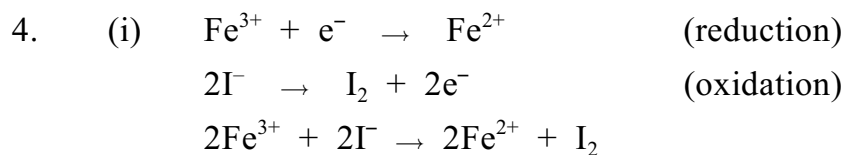
(iii) its oxidation number decreases.

e.g. the oxidation number of Cu in both (i) and (ii) decreases from +II on the LHS to zero on the RHS.

3. (i) +I The O.N. of a simple cation = the charge on that cation.
 (ii) 0 The O.N. of an element = zero.
 (iii) +VI The sum of the O.N. of all the atoms = charge on the ion = -II.
 The sum of the O.N. of the four O atoms = $4 \times -II = -VIII$.
 \therefore O.N. of the S atom = +VI.

- (iv) +II The Fe in this compound must be Fe^{2+} ion to balance the SO_4^{2-} ionic charge, so its O.N. = +II.
- (v) 0 The O.N. of an element = zero.
- (vi) +I The Na in this compound must be Na^+ to balance the Cl^- ionic charge, so its O.N. = +I [Na, like all Group I elements, only has the +I oxidation state.]
- (vii) +I The Na in this compound must be Na^+ to balance the SO_4^{2-} ionic charge, so its O.N. = +I.
- (viii) +V The PO_4^{3-} ionic charge = sum of the O.N. of the P and the four O atoms. The O.N. of each O atom = -II, so O.N. of the P atom = +V.
- (ix) +IV The CO_3^{2-} ionic charge = sum of the O.N. of the C and the three O atoms. The O.N. of each O atom = -II, so O.N. of the C atom = +IV.
- (x) +II The Mg in this compound must be Mg^{2+} to balance the O^{2-} ionic charge, so its O.N. = +II.
- (xi) +II The Mg in this compound must be Mg^{2+} to balance the two Cl^- ionic charges, so its O.N. = +II. [Mg like the other Group II elements only has the +II oxidation state.]
- (xii) -III Hydrogen normally has the +I oxidation state except in ionic hydrides. The sum of the O.N. of the atoms = zero, so the N atom must have O.N. = -III.
- (xiii) +V The sum of the O.N. of the atoms = -I and the O atoms each have O.N. = -II so the O.N. of the N atom = +V.
- (xiv) -III The O.N. of a simple anion = charge on that anion = -III
- (xv) +II The Cu in this compound must be Cu^{2+} to balance the O^{2-} ionic charge, so its O.N. = +II.
- (xvi) +I The Cu in this compound must be Cu^+ to balance the O^{2-} ionic charge, so its O.N. = +I.
- (xvii) +III The Fe in this compound must be Fe^{3+} to balance the total O^{2-} ionic charge, so its O.N. = +III.
- (xviii) +VI The $\text{Cr}_2\text{O}_7^{2-}$ ionic charge = sum of the O.N. of the two Cr atoms and the seven O atoms. The O.N. of each O atom = -II, so O.N. of each Cr atom = +VI.

- (xix) +III The O.N. of a simple cation = charge on that cation = +III.
 (xx) +IV The sum of the O.N. of the C and the two O atoms = zero and the O.N. of each O atom = -II so O.N. of the C atom = +IV.



5. (a) Reducing agent: Cl^- Oxidizing agent: MnO_2
 Atom oxidized: Cl as Cl^- (O.N. = -I) to Cl in Cl_2 (O.N. = 0).
 Atom reduced: Mn in MnO_2 (O.N. = +IV) to Mn as Mn^{2+} (O.N. = +II).

(b) Reducing agent: SO_2 Oxidizing agent: NO_3^-
 Atom oxidized: S in SO_2 (O.N. = +IV) to S in SO_4^{2-} (O.N. = +VI).
 Atom reduced: N in NO_3^- (O.N. = +V) to N in NO (O.N. = +II).

(c) Reducing agent: Ag Oxidizing agent: H_2S
 Atom oxidized: Ag as the element (O.N. = 0) to Ag^+ in Ag_2S (O.N. = +I).
 Atom reduced: H in H_2S (O.N. = +I) to H in H_2 (O.N. = 0).

(d) Reducing agent: H_2S Oxidizing agent: NO_3^-

Atom oxidized: S in H_2S (O.N. = -II) to S as the element (O.N. = 0).

Atom reduced: N in NO_3^- (O.N. = +V) to N in NO_2 (O.N. = +IV).

(e) Reducing agent: Pb Oxidizing agent: PbO_2

Atom oxidized: Pb as the element (O.N. = 0) to Pb as Pb^{2+} (O.N. = +II).

Atom reduced: Pb in PbO_2 (O.N. = +IV) to Pb^{2+} (O.N. = +II).

(This example shows that two different oxidation states of the same element can lead to both oxidation and reduction of the one type of atom; this reaction is the basis for the lead acid battery).

(f) Reducing agent: Fe^{2+} Oxidizing agent: $\text{Cr}_2\text{O}_7^{2-}$

Atom oxidized: Fe as Fe^{2+} (O.N. = +II) to Fe as Fe^{3+} (O.N. = +III).

Atom reduced: Cr in $\text{Cr}_2\text{O}_7^{2-}$ (O.N. = +VI) to Cr as Cr^{3+} (O.N. = +III).

(g) Reducing agent: Br^- Oxidizing agent: Cl_2

Atom oxidized: Br as Br^- (O.N. = -I) to Br as the element Br_2 (O.N. = 0).

Atom reduced: Cl as the element Cl_2 (O.N. = 0) to Cl as Cl^- (O.N. = -I).

(h) Reducing agent: H_2 Oxidizing agent: CuO

Atom oxidized: H as the element H_2 (O.N. = 0) to H in H_2O (O.N. = +I).

Atom reduced: Cu in CuO (O.N. = +II) to Cu as the element (O.N. = 0).

(i) Reducing agent: Sn^{2+} Oxidizing agent: H_3AsO_4

Atom oxidized: Sn as Sn^{2+} (O.N. = +II) to Sn as Sn^{4+} (O.N. = +IV).

Atom reduced: As in H_3AsO_4 (O.N. = +V) to As in H_3AsO_3 (O.N. = +III).

(j) Reducing agent: Fe^{2+} Oxidizing agent: MnO_4^-

Atom oxidized: Fe as Fe^{2+} (O.N. = +II) to Fe as Fe^{3+} (O.N. = +III).

Atom reduced: Mn in MnO_4^- (O.N. = +VII) to Mn as Mn^{2+} (O.N. = +II).

6. NiO +II Ni is present as the Ni^{2+} ion

Na_2 S -II S is present as the S^{2-} ion

Mn O_4^- +VII The sum of the O.N. of the Mn atom and the four O atoms = charge on the ion = -I. Given each O atom has O.N. = -II, then the Mn atom has O.N. = +VII

K_2 Cr O_4 +VI The chromate ion has the formula CrO_4^{2-} . The sum of the O.N. of the Cr atom and the four O atoms = charge on the ion = -II. Given each O atom has O.N. = -II, then the Cr atom has O.N. = +VI.

H NO_3 +V The nitrate ion has the formula NO_3^- . The sum of the O.N. of the N atom and the three O atoms = charge on the ion = -I. Given each O atom has O.N. = -II, then the N atom has O.N. = +V.

$\underline{\text{N}}\text{O}_2$ +IV The sum of the O.N. of the N and the two O atoms = zero. As the O.N. of each O atom = -II, then the O.N. of the N atom = +IV.

$\underline{\text{S}}\text{O}_3$ +VI The sum of the O.N. of the S and the three O atoms = zero. As the O.N. of each O atom = -II, then the O.N. of the S atom = +VI.

$\underline{\text{H}_3}\underline{\text{P}}\text{O}_4$ +V The phosphate ion has the formula PO_4^{3-} . The sum of the O.N. of the P atom and the four O atoms = charge on the ion = -III. Given each O atom has O.N. = -II, then the P atom has O.N. = +V.

$\underline{\text{H}_2}$ 0 All elements have O.N. = zero.

$\underline{\text{C}}\text{O}_4^-$ +VII The sum of the O.N. of the Cl atom and the four O atoms = charge on the ion = -I. Given each O atom has O.N. = -II, then the Cl atom has O.N. = +VII.

7. (a) not redox (no atom changes its oxidation state).

(b) not redox (no atom changes its oxidation state).

(c) redox reaction

Mg, the element (O.N. = 0) is oxidised to Mg^{2+} which has O.N. = +II.

The H^+ ion (O.N. = +I) is reduced to the element, H_2 , which has O.N. = 0.

(d) not redox (no atom changes its oxidation state).

(e) redox reaction.

Cu is oxidised from O.N. = 0 in the element to O.N. = +II in CuS.

S is reduced from O.N. = 0 in the element to O.N. = -II in CuS.

8. **Oxidation Is Loss Reduction Is Gain (of electrons) OILRIG**

9. The oxidation number of S in both SO_2 and SO_3^{2-} is +IV. Oxygen atoms in both species all have O.N. = -II. As there has been no change in the oxidation number of any atom, this is not a redox reaction.

10. The oxidation state (oxidation number) of S in these species is

SO_4^{2-} (+VI) SO_3^{2-} (+IV) S(s) (0) SO_2 (+IV) H_2S (-II)

(derived using the O.N. of O atoms = -II and H atoms = +I).

Thus S has its highest oxidation number in SO_4^{2-} and its lowest in H_2S .

11. The oxidation state (oxidation number) of N in these species is

NO_2 (+IV) NO (+II) HNO_3 (+V) N_2 (0)

HNO_2 (+III) N_2O (+I) Na_3N (-III)

Thus N has its highest oxidation number in HNO_3 and its lowest in Na_3N .