Cosmo School

Chemistry

Class-10

Chapter-7

Chemical reactions

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Revision worksheet with lecture sheet-3

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Unit-1: Oxidation number or oxidation state

Oxidation number is a number assigned to an element in chemical combination which represents the number of electrons lost (if the number is positive) or gained (if the number is negative), by an atom of that element in the compound. The oxidation number, sometimes referred to as oxidation state.

There are the rules that are used to figure out oxidation numbers.

- The first rule is this, an element by itself always has an oxidation number of 0. It means that there are a lot of chemical substances that have just one element that element is not combined with any other elements. So the oxidation state of any pure element is always zero. So the oxidation state of oxygen gas(O_2) as a pure element is zero, fluorine gas(F_2) as the pure element is zero, even phosphorus(P or P_4) as a pure element is zero. So, there is no charges and it's only one pure element and it is not a compound. the oxidation state will always be zero.
- The other rule is about monatomic ions. These are ions that are made of only one and so like this for monatomic ions their oxidation number is the same as their ion charge. So for K^+ here it's oxidation number is going to be +1. For N^{3-} ion, it will have an oxidation number of -3 and for Mg^{2+}, here is going to have an oxidation number of +2.
- When we write oxidation numbers, we write this the sign first so plus(+) minus (-) and then the number after. This is the opposite of how we write ionic charges. So just keep that in mind the charge might be 2+ on magnesium but the oxidation number is +2.
- There is another example this is the peroxide ion(O_2^{2-}) to find the oxidation state of each oxygen atom. In this ion you can write an equation is two oxygen atoms with the total charge of -2. So individually each oxygen atom has a charge of -1. So that is the oxidation state of oxygen individually in the peroxide ion. In the superoxide ion(O_2^-) if you want to find the oxidation state, you need to divide the total charge by 2. So each oxygen atom has a net charge of - ½. So two of them combined will have a net charge of -1. Whenever you have fluorine inside a compound when it's not a pure element fluorine is always going to have a negative 1(-1) oxidation state. Fluorine is the most electronegative element.

For O_2^{2-}, $2O = -2$ For O_2^{-}, $2O = -1$

$O = -1$ $O = -1/2$ or 0.5

- When oxygen is in a compound, it's going to have a -2 oxidation state unless it's bonded to fluorine or unless you hear the name peroxide or superoxide. Whenever you hear the name peroxide oxygen has a -1 oxidation state if you hear the word superoxide it has a -1/2 oxidation state if you hear the word oxide then the oxidation state is -2.
- Now hydrogen will have an oxidation state of +1 when bonded to a nonmetal. When bonded to a metal, hydrogen will have an oxidation state of -1 and really the key is electronegativity. Hydrogen is more electronegative than most metals. That's why it bears a negative charge but hydrogen is usually less electronegative than most nonmetals and so that's why there's a positive charge. So typically, the element that is more electronegative is the one that usually carries the negative charge.
- Now let's work on some examples.
  What is the oxidation state of magnesium and chlorine in $MgCl_2$ compound? By the way most halogens are usually -1, Chlorine technically has a -1 charge like fluorine. If we write an equation,
  $$Mg + 2\,Cl = 0$$
  This whole compound is neutral so therefore the total charge is zero. Now if chlorine has a -1 oxidation state that means magnesium has to have a +2 oxidation state. You can literally solve it and it makes sense magnesium is an alkaline earth metal which typically has a +2 charge.
- You could solve another example, find the oxidation state of vanadium and oxygen in a compound, V_2O_5. This is called vanadium oxide.
  So whenever you hear the word oxide, oxygen has a ………… charge. So we got two vanadium atoms with ………… oxygen atoms with a net charge of zero. So each oxygen atom has an oxidation state of -2.
- Some examples containing polyatomic ions, consider sulfate(SO_4^{2-}).
  What is the oxidation state of sulfur in sulfate?
  We know oxygen is usually -2. So let's write an equation sulfur plus 4 oxygen atoms has a net charge of -2.
  $$S + 4O = -2$$
  Suppose, the oxidation no. of S = x

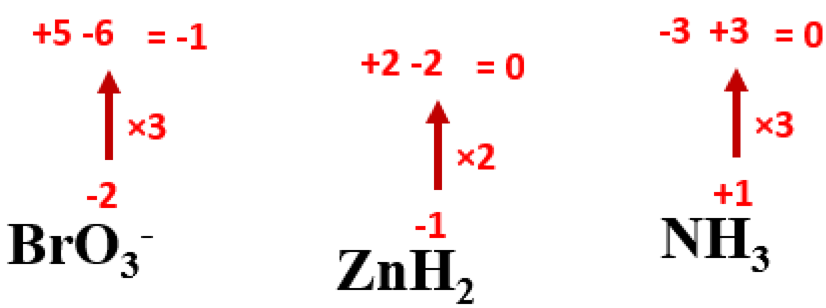
  The oxidation state of sulfur in sulfate will be …………

- Now two more examples, $BrCl_3$ and IBr_5. Find the oxidation state of every element in these examples.
  So most halogens like fluorine, chlorine, bromine and iodine, they typically have a -1 charge. But in $BrCl_3$, both bromine and chlorine can't be negative. So which one is negative and which one is positive? Keep in mind bromine has an electronegativity value

of 2.8, chlorine is 3.0 and iodine is 2.5. So in this example chlorine bears the partial charge and bromine is partially positive. So, therefore, chlorine is going to have its natural oxidation state of -1 and for bromine we need to calculate it.
So it's going to be Br + 3 Cl and that's equal to 0. So this is going to be 3 times -1 and so we can see that bromine has an oxidation state of +3.

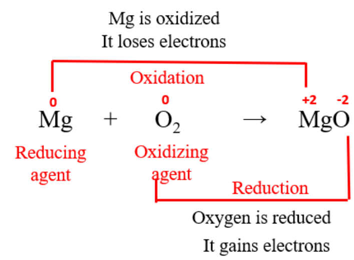
- Now in the second example, IBr_5, bromine is going to carry the partial ………….charge and iodine carries the partial …………….. charge. In any compound, usually the electropositive element is written first the electronegative element is there in second. So the one that you see on the right side is usually the one that carries the natural charge. So in this case bromine is going to have its natural oxidation state of …………
So iodine is going to have an oxidation state of ………… in IBr_5.

Exercise-1:



1. Explain the three processes used to calculate the oxidation number shown in figure above.
2. What is oxidation number (oxidation state)?
3. Mention the differences between valency and oxidation number.
4. Could the oxidation number of any element be a whole number or fraction? Explain according to the compounds $K_2Cr_2O_7$ and $Na_2S_4O_6$.
5. Determine the oxidation number of central atoms in the following compounds.
6. $Na_2S_2O_3$, H_2SO_4, $KMnO_4$, H_3PO_4, HNO_3 and Na_2CO_3

Unit-2:
Example-1: Redox reaction



Magnesium reacts with oxygen gas produces magnesium oxide in an oxidation-reduction reaction or simply a redox reaction. Electrons are being transferred from one element to another. Since the oxidation number of any pure element is zero, The oxidation states of Mg and O_2 is 0. Now in magnesium oxide(MgO), magnesium being an alkaline earth metal has a charge of +2 and oxygen has a charge of -2. Now notice that the oxidation state of magnesium went from 0 to +2. So, the oxidation state increased. Whenever the oxidation state goes up, oxidation is taking place, the substance is said to be oxidized. In the case of oxygen, the oxidation number decreased from 0 to -2.
Whenever the oxidation number decreases that means goes down, the substance is being reduced. Oxidation always occurs with a loss of electrons and reduction is associated with a gain of electrons. Metals, they like to give away electrons, they like to form metal cations as they give away electrons, they will acquire a positive charge.
Nonmetals like oxygen, they like to acquire electrons and so they will develop a negative charge. The substance that is oxidized is known as to reducing agent or reductant and the substance that is reduced is known as the oxidizing agent or oxidant. Metals are reducing agents because they will cause the other substance to be reduced.
Nonmetals like oxygen gas, fluorine, they are oxidizing agents because they cause another substance to be oxidized.

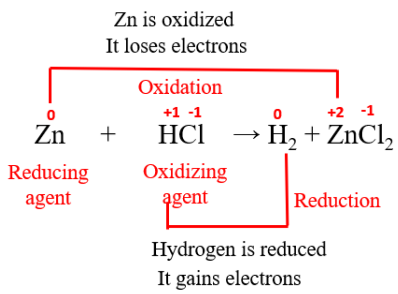
Now magnesium (Mg) is changing into the magnesium ion (Mg^{+2}). This magnesium ion has a +2 charge and in order to become +2 cation, it has to lose two electrons. Now because magnesium lost electrons, this reaction is known as a half reaction. That is the oxidation part of the reaction. When a substance loses electrons, it is being oxidized.

Now oxygen, I'm going to write atomic oxygen because this reaction is not balanced. Individually oxygen acquires two electrons and turns into oxide(O^{-2}). This second half reaction is the reduction part, since oxygen is acquiring electrons. Whenever a substance gained electrons or if the oxidation number decreases, its reduction. Now for half reactions, anytime you have electrons on the right side, it is going to be an oxidation half-reaction. Whenever the electrons are on the left side, it is a reduction half-reaction.

$$Mg \rightarrow Mg^{+2} + 2e^- \quad \text{[Oxidation half reaction]}$$

$$O + 2e^- \rightarrow O^{-2} \quad \text{[Reduction half reaction]}$$

Example-2: Single displacement or substitution reaction

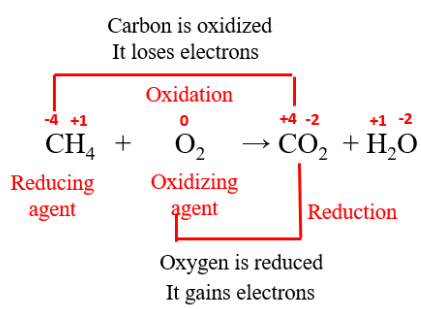


$Zn \rightarrow Zn^{+2} + 2e^-$ [Oxidation half reaction]

$H^{+1} + e^- \rightarrow H$ [Reduction half reaction]

zinc metal reacts with hydrochloric acid to produce hydrogen gas and zinc chloride. Identify the substance that is oxidized and the substance that is reduced and also identify the oxidizing agent and a reducing agent. Now if you get a question like this on a test to find a substance that is oxidized or reduced. Always look at the reactants, both are present in reactants and don't look at the products. Now let's find the oxidation states of everything. The oxidation state of any pure element is always zero. Now whenever hydrogen is bonded to a nonmetal, it's going to have a +1 oxidation state which means chlorine has to be -1. halogens are usually -1 except with Oxygen, so Cl still -1 which means Zn has to have a +2 charge. Now the oxidation of zinc changes from 0 to +2. So zinc is being oxidized. Hydrogen changes from +1 to 0. So HCl as a substance is being reduced, even though only the hydrogen portion of that substance is being reduced.

Example-3: Combustion reaction



- So identify the oxidation states of every element in this reaction and then find the substance that is oxidized reduced and identify the oxidizing agent and a reducing agent.

Methane reacts with oxygen gas and it produces carbon dioxide and water.
In the above reaction, the pure elements which is only oxygen gas. The oxidation state of O_2 is going to be zero. Now whenever hydrogen is bonded to a nonmetal, it is going to have a +1 oxidation state. And when oxygen is found in a compound, the oxidation state is -2 except when it is bonded to fluorine or except in peroxides or superoxides, its different. In methane(CH_4) we have one carbon atom and four hydrogen atoms which has to add up to make zero because methane is neutral in charge. Now each hydrogen atom has an oxidation state of +1. So therefore carbon has to have an oxidation state of -4 in methane.

In CO_2, 1 Carbon & 2 oxygen atoms which equals in the charge of zero and each oxygen has a charge of -2. So in this case, carbon is going to have an oxidation state of +4. So carbon changes from - 4 to +4 .Ttherefore the oxidation number of carbon is increasing which means carbon is oxidized or that means methane is oxidized. Now hydrogen doesn't change but oxygen changes from zero to -2. So therefore, for oxygen gas, the oxidation number is decreasing. So it's being reduced which means that methane is the reducing agent, it helps to reduce O_2 and oxygen gas is being reduced which makes it the oxidizing agent. The whole substance methane as being oxidized because carbon is part of methane.

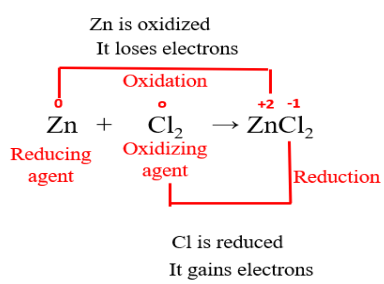
> Now if you don't want to look for the oxidation numbers, there is a quick way to tell if the reaction is going to be a redox reaction. if you see a pure element on one side and then that element being part of a compound on the other side, it's always going to be a redox reaction. there's a transfer of electrons.

Example-4 & 5: Addition reactions

Now let's take more two examples and determine if it's a redox reaction or not. so both of these examples are synthesis reactions also known as combination or addition reactions. sometimes a combination reaction can be a redox reaction and sometimes it won't be. now if you look at the first example, there are no pure elements, all we have is compounds. When you see that type of change, chances are, it's not a redox reaction. but if you look at the second example we have a pure element zinc and then that same element is in a compound in product side. so the second example is a redox reaction and let's prove it using oxidation numbers.

$$\overset{+2}{Mg}\overset{-2}{O} + \overset{+1}{H_2}\overset{-2}{O} \rightarrow \overset{+2}{Mg}(\overset{-2}{O}\overset{+1}{H})_2$$

So let's start with the first example, oxygen has a -2 charge, magnesium in magnesium oxide has a +2 charge, hydrogen has a +1 charge on left side. Now magnesium hydroxide is composed of Mg and two hydroxide ions but magnesium still has an oxidation state of +2. And in magnesium hydroxide, oxygen is still -2 and hydrogen is +1. So the oxidation state does not change for any element that means there was no transfer of electrons and so this is not a redox reaction.



Now for the second example zinc and chlorine has an oxidation state of zero on left side and in zinc chloride, Cl has -1 oxidation state. zinc has a +2 oxidation state. So zinc is being oxidized, therefore, there is a transfer of electrons which makes this reaction a redox reaction.

Example-6 & 7: Decomposotion reactions

Now consider the two decomposition reactions. When mercury oxide is heated, it decomposes into mercury and oxygen gas and if we heat calcium carbonate, it will decompose into calcium oxide and carbon dioxide gas. So which of these decomposition reactions, is the redox reaction?

$$HgO \rightarrow Hg + O_2$$ Redox reaction

Pure element

$$CaCO_3 \rightarrow CaO + CO_2$$ Non-redox reaction

Looking at the first example, we have oxygen as a pure element on the right side and then that same element is within a compound. so the first example is a redox reaction.

In the second example, there are no pure elements, all we have are just compounds. so the second example is not a redox reaction. That is a quick way to tell if it's a redox or not.

- All combustion reactions are redox reactions.
- All single displacement reactions are redox reactions.
- Now synthesis and combination reactions sometimes the redox sometimes or not and the same is true for decomposition reactions.
- Double replacement reactions are never redox reactions so that includes acid-base reactions, precipitation reactions and other types of double replacement reactions.
- And a quick way to tell is if you see a pure element on one side and then the same element within a compound on the other side, it's going to be a redox reaction.

Exercise-2:

1. What is oxidant (oxidizing agent) and reductant (reducing agent)?

2. Determine which ones are redox and which ones are not? Analyze using oxidation number.
   i) $C_2H_5OH + O_2 \rightarrow CO_2 + H_2O$
   ii) $NaOH + HCl \rightarrow H_2O + NaCl$
   iii) $CO_2 + H_2O \rightarrow H_2CO_3$
   iv) $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
   v) $Al + CuCl_2 \rightarrow Cu + AlCl_3$
   vi) $AlCl_3(s) + 3H_2O \rightarrow Al(OH)_3(s) + 3HCl(aq)$
   vii) $FeCl_3 + 3H_2O \rightarrow Fe(OH)_3 + 3HCl$
   viii) $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + NaCl(aq)$
   ix) $FeCl_2(aq) + Cl_2(g) \rightarrow FeCl_3(aq)$