

Chapter 7: Redox Reactions

NCERT · Class 11 · Chemistry

Questions

Question 1

Assign oxidation number to the underlined elements in each of the following species:

(a)



(b)



(c)



(d)



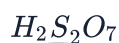
(e)



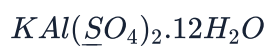
(f)



(g)

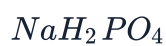


(h)



Solution

(a)



Let x be the oxidation no. of P.

Oxidation no. of Na = +1

Oxidation no. of H = +1

Oxidation no. of O = -2



Then,

$$1(+1) + 2(+1) + 1(x) + 4(-2) = 0$$

$$1 + 2 + x - 8 = 0$$

$$x = +5$$

Therefore, oxidation no. of P is +5.

(b)



Let x be the oxidation no. of S.

Oxidation no. of Na = +1

Oxidation no. of H = +1

Oxidation no. of O = -2

Then,

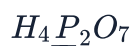
$$1(+1) + 1(+1) + 1(x) + 4(-2) = 0$$

$$1 + 1 + x - 8 = 0$$

$$x = +6$$

Therefore, oxidation no. of S is +6.

(c)



Let x be the oxidation no. of P.

Oxidation no. of H = +1

Oxidation no. of O = -2

Then,

$$4(+1) + 2(x) + 7(-2) = 0$$

$$4 + 2x - 14 = 0$$

$$2x = +10$$

$$x = +5$$

Therefore, oxidation no. of P is +5.

(d)





Let x be the oxidation no. of Mn.

Oxidation no. of K = +1

Oxidation no. of O = -2

Then,

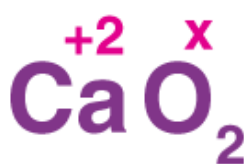
$$2(+1) + x + 4(-2) = 0$$

$$2 + x - 8 = 0$$

$$x = +6$$

Therefore, oxidation no. of Mn is +6.

(e)



Let x be the oxidation no. of O.

Oxidation no. of Ca = +2

Then,

$$(+2) + 2(x) = 0$$

$$2 + 2x = 0$$

$$2x = -2$$

$$x = -1$$

Therefore, oxidation no. of O is -1.

(f)



Let x be the oxidation no. of B.

Oxidation no. of Na = +1

Oxidation no. of H = -1

Then,

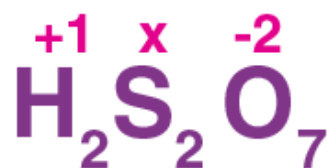
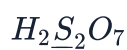
$$1(+1) + 1(x) + 4(-1) = 0$$

$$1 + x - 4 = 0$$

$$x = +3$$

Therefore, oxidation no. of B is +3.

(g)



Let x be the oxidation no. of S.

Oxidation no. of H = +1

Oxidation no. of O = -2

Then,

$$2(+1) + 2(x) + 7(-2) = 0$$

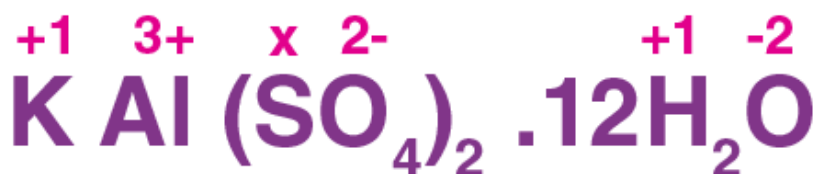
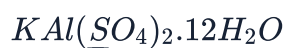
$$2 + 2x - 14 = 0$$

$$2x = +12$$

$$x = +6$$

Therefore, oxidation no. of S is +6.

(h)



Let x be the oxidation no. of S.

Oxidation no. of K = +1

Oxidation no. of Al = +3

Oxidation no. of O = -2

Oxidation no. of H = +1

Then,

$$1(+1) + 1(+3) + 2(x) + 8(-2) + 24(+1) + 12(-2) = 0$$

$$1 + 3 + 2x - 16 + 24 - 24 = 0$$

$$2x = +12$$

$$x = +6$$

Therefore, oxidation no. of S is +6.

OR

Ignore the water molecules because it is neutral. Then, the summation of the oxidation no. of all atoms of water molecules can be taken as 0. Hence, ignore the water molecule.

$$1(+1) + 1(+3) + 2(x) + 8(-2) = 0$$

$$1 + 3 + 2x - 16 = 0$$

$$2x = 12$$

$$x = +6$$

Therefore, oxidation no. of S is +6.

Question 2

What are the oxidation numbers of the underlined elements in each of the following and how do you rationalise your results?

(a)



(b)



(c)



(d)



(e)



Solution

(a)



Let x be the oxidation no. of I.

Oxidation no. of K = +1

Then,

$$1(+1) + 3(x) = 0$$

$$1 + 3x = 0$$

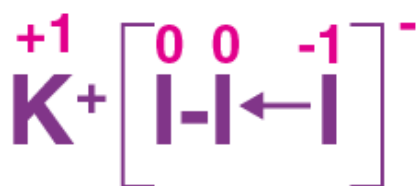
$$x =$$

$$-\frac{1}{3}$$

Oxidation no. cannot be fractional. Hence, consider the structure of



In



Therefore, in



(b)



Let x be the oxidation no. of S.

Oxidation no. of H = +1

Oxidation no. of O = -2

Then,

$$2(+1) + 4(x) + 6(-2) = 0$$

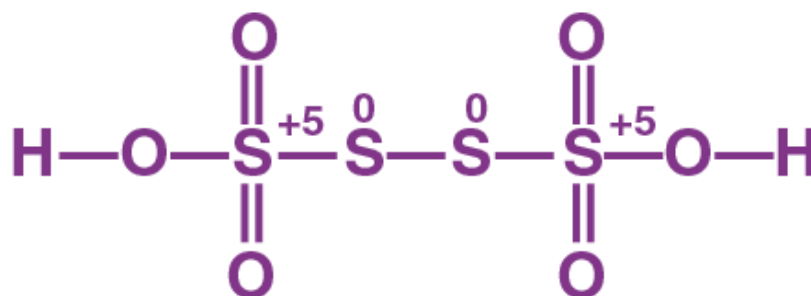
$$2 + 4x - 12 = 0$$

$$4x = 10$$

$$x =$$

$$+2\frac{1}{2}$$

Oxidation no. cannot be fractional. Therefore, S would be present with different oxidation states in the molecule.



The oxidation no. of two out of the four S atoms is +5 while that of the other two atoms is 0.

(c)



Let x be the oxidation no. of Fe.

Oxidation no. of O = -2

Then,

$$3(x) + 4(-2) = 0$$

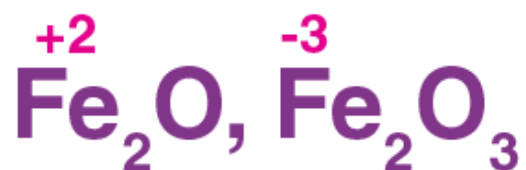
$$3x - 8 = 0$$

$$x =$$

$$\frac{8}{3}$$

Oxidation no. cannot be fractional.

One of the three atoms of Fe has oxidation no. +2 and the other two atoms of Fe have oxidation no. +3.



(d)



Let x be the oxidation no. of C.

Oxidation no. of O = -2

Oxidation no. of H = +1

Then,

$$2(x) + 4(+1) + 1(-2) = 0$$

$$2x + 4 - 2 = 0$$

$$x = -2$$

Therefore, oxidation no. of C is -2.

(e)





Let x be the oxidation no. of C.

Oxidation no. of O = -2

Oxidation no. of H = +1

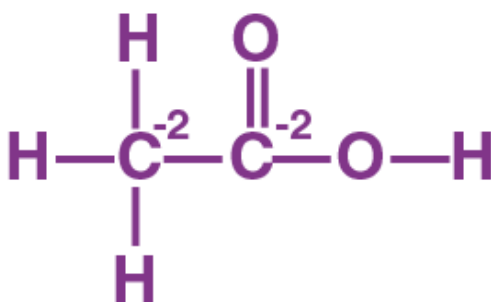
Then,

$$2(x) + 4(+1) + 2(-2) = 0$$

$$2x + 4 - 4 = 0$$

$$x = 0$$

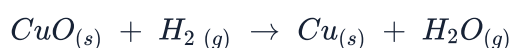
Therefore, the average oxidation no. of C is 0. Both the carbon atoms are present in different environments, so they cannot have the same oxidation no. Therefore, carbon has oxidation no. of +2 and -2 in



Question 3

Justify that the following reactions are redox reactions:

(a)



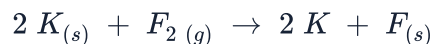
(b)



(c)



(d)

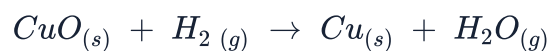


(e)



Solution

(a)



Oxidation no. of Cu and O in



Oxidation no. of



Oxidation no. of Cu is 0.

Oxidation no. of H and O in



The oxidation no. of Cu decreased from +2 in



The oxidation no. of H increased from 0 to +1 in



Therefore, the reaction is a redox reaction.

(b)



In the above reaction,

Oxidation no. of Fe and O in



Oxidation no. of C and O in CO is +2 and -2, respectively.

Oxidation no. of Fe is 0.

Oxidation no. of C and O in



The oxidation no. of Fe decreased from +3 in



The oxidation no. of C increased from 0 to +2 in CO to +4 in



Therefore, the reaction is a redox reaction.

(c)



the above reaction,

Oxidation no. of B and Cl in



Oxidation no. of Li, Al and H in



Oxidation no. of B and H in



Oxidation no. of Li and Cl in LiCl is +1 and -1, respectively.

Oxidation no. of Al and Cl in



The oxidation no. of B decreased from +3 in

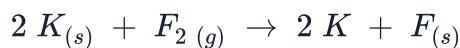


The oxidation no. of H increased from -1 in



Therefore, the reaction is a redox reaction.

(d)



In the above reaction,

Oxidation no. of K is 0.

Oxidation no. of F is 0.

Oxidation no. of K and F in KF is +1 and -1, respectively.

The oxidation no. of K increased from 0 in K to +1 in KF. That is K is oxidized to KF.

The oxidation no. of F decreased from 0 in



Therefore, the reaction is a redox reaction.

(e)



In the above reaction,

Oxidation no. of N and H in



Oxidation no. of



Oxidation no. of N and O in NO is +2 and -2, respectively.

Oxidation no. of H and O in



The oxidation no. of N increased from -3 in



The oxidation no. of





Therefore, the reaction is a redox reaction.

Question 4

Fluorine reacts with ice and results in the change:



Justify that this reaction is a redox reaction

Solution



In the above reaction,

Oxidation no. of H and O in



Oxidation no. of



Oxidation no. of H and F in HF is +1 and -1, respectively.

Oxidation no. of H, O and F in HOF is +1, -2 and +1, respectively.

The oxidation no. of F increased from 0 in



The oxidation no. of F decreased from 0 in



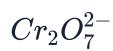
Therefore, F is both reduced as well as oxidized. So, it is a redox reaction.

Question 5

Calculate the oxidation no. of sulphur, chromium and nitrogen in



,



and



. Suggest structure of these compounds. Count for the fallacy.

Solution

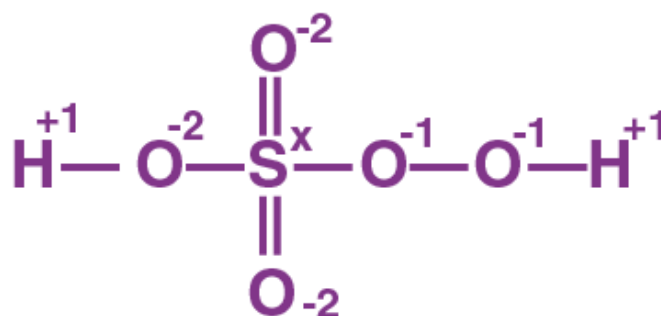
For



Let x be the oxidation no. of S.

Oxidation no. of O = -2

Oxidation no. of H = +1



Then,

$$2(+1) + 1(x) + 5(-2) = 0$$

$$2 + x - 10 = 0$$

$$x = +8$$

But the oxidation no. of S cannot be +8 as S has 6 valence electrons. Therefore, the oxidation no. of S cannot be more than +6.

The structure of



Now,

$$2(+1) + 1(x) + 3(-2) + 2(-1) = 0$$

$$2 + x - 6 - 2 = 0$$

$$x = +6$$

Therefore, the oxidation no. of S is +6.

For



Let x be the oxidation no. of Cr.

Oxidation no. of O = -2

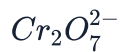
Then,

$$2(x) + 7(-2) = -2$$

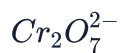
$$2x - 14 = -2$$

$$x = +6$$

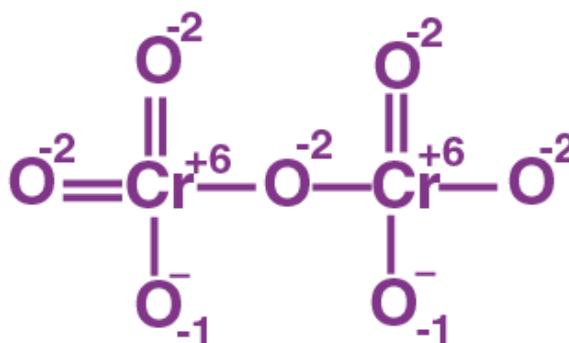
There is no fallacy about the oxidation no. of Cr in



The structure of



Each of the two Cr atoms has the oxidation no. of +6.



For



Let x be the oxidation no. of N.

Oxidation no. of O = -2

Then,

$$1(x) + 3(-2) = -1$$

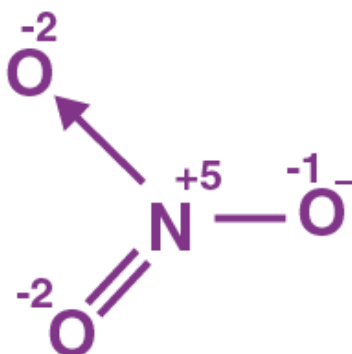
$$x - 6 = -1$$

$$x = +5$$

There is no fallacy about the oxidation no. of N in



The structure of



Nitrogen atom has the oxidation no. of +5.

Question 6

Write formulas for the following compounds:

(a) Mercury (II) chloride

(b) Nickel (II) sulphate

(c) Tin (IV) oxide

(d) Thallium (I) sulphate

(e) Iron (III) sulphate

(f) Chromium (III) oxide

Solution

(a) Mercury (II) chloride



(b) Nickel (II) sulphate



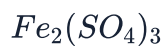
(c) Tin (IV) oxide



(d) Thallium (I) sulphate



(e) Iron (III) sulphate



(f) Chromium (III) oxide



Question 7

Suggest a list of the substances where carbon can exhibit oxidation states from -4 to +4 and nitrogen from -3 to +5.

Solution

The compound where carbon has oxidation no. from -4 to +4 is as given below in the table:

Compounds	Oxidation no. of carbon
CH_2Cl_2	0
$HC \equiv CH$	-1
$ClC \equiv CCl$	+1
CH_3Cl	-2

$CHCl_3$, CO	+2
$H_3C - CH_3$	-3
$Cl_3C - CCl_3$	+3
CH_4	-4
CCl_4 , CO_2	+4

Compounds	Oxidation no. of nitrogen
N_2	0
N_2H_2	-1
N_2O	+1
N_2H_4	-2
NO	+2
NH_3	-3
N_2O_3	+3
NO_2	+4
N_2O_5	+5

Question 8

While sulphur dioxide and hydrogen peroxide can act as oxidising as well as reducing agents in their reactions, ozone and nitric acid act only as oxidants. Why?

Solution

In sulphur dioxide (



Hence,



In hydrogen peroxide (



Therefore,



In ozone (



Therefore,



In nitric acid (



Therefore,



Question 9

Consider the reactions:

(a)



(b)



Why it is more appropriate to write these reactions as :

(a)



(b)



Also suggest a technique to investigate the path of the above (a) and (b) redox reactions

Solution

(a)

Step 1 :

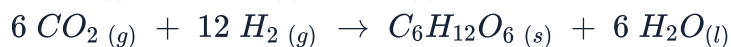


Step 2 :

The



The net reaction is as given below:



This is the suitable way to write the reaction as the reaction also produces water molecules in the photosynthesis process.

The path can be found with the help of radioactive



(b)

Step 1 :





Step 2 :



The path can be found with the help of



Question 10

The compound AgF_2 is unstable compound. However, if formed, the compound acts as a very strong oxidising agent. Why?

Solution

The oxidation no. of Ag in



Question 11

Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. Justify this statement giving three illustrations.

Justify the above statement with three examples.

Solution

When there is a reaction between a reducing agent and an oxidizing agent, a compound is formed, which has a lower oxidation number if the reducing agent is in excess. A compound is formed which has a higher oxidation number if the oxidizing agent is in excess.

(i)



and



are reducing and oxidizing agent respectively.

In an excess amount of



is reacted with



, then



would be produced, where the oxidation no. of P is +3.



If



is reacted with excess of



, then



would be produced, where the oxidation no. of P is +5.



(ii) K and



acts as a reducing agent and oxidizing agent respectively.

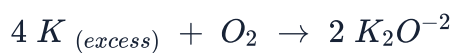
If an excess of K reacts with



, it produces



. Here, the oxidation number of O is -2.



If K reacts with an excess of



, it produces



, where the oxidation number of O is -1.



(iii) C and



acts as a reducing agent and oxidizing agent respectively.

If an excess amount of C is reacted with an insufficient amount of



, then it produces CO, where the oxidation number of C is +2.



If C is burnt in excess amount of



, then



is produced, where the oxidation number of C is +4.



Question 12

How do you count for the following observations?

(a) Though alkaline potassium permanganate and acidic potassium permanganate both are used as oxidants, yet in the manufacture of benzoic acid from toluene we use alcoholic potassium permanganate as an oxidant. Why? Write a balanced redox equation for the reaction.

(b) When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smelling gas HCl , but if the mixture contains bromide then we get red vapour of bromine. Why?

Solution

(a) While manufacturing benzoic acid from toluene, alcoholic potassium permanganate is used as an oxidant due to the given reasons.

(i) In a neutral medium,



ions are produced in the reaction. Due to that, the cost of adding an acid or a base can be reduced.

(ii)

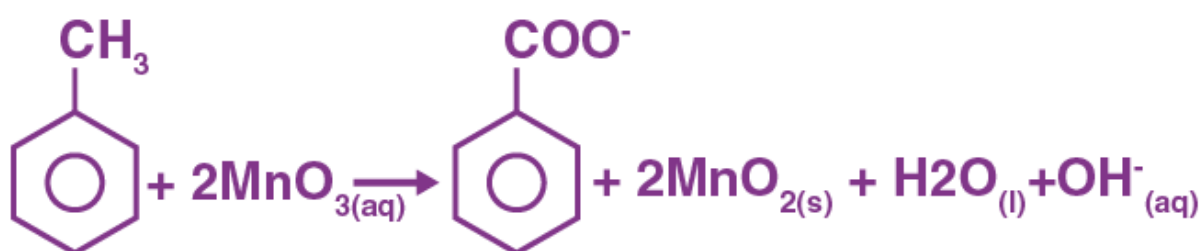


and alcohol are homogeneous to each other as they are polar. Alcohol and toluene are homogeneous to each other because both are organic compounds. Reactions can proceed at a faster rate in a homogeneous medium compared to a heterogeneous medium. Therefore, in alcohol,



and toluene can react at a faster rate.

The redox reaction is as given below:



(b) When concentrated



is added to an inorganic mixture containing bromide, firstly HBr is produced. HBr , a strong reducing agent, reduces



to



with the evolution of bromine's red vapour.



When concentrated



I added to an inorganic mixture containing chloride, a pungent smelling gas (HCl) is evolved. HCl, a weak reducing agent, cannot reduce



to



Question 13

Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions:

(a)



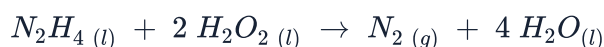
(b)



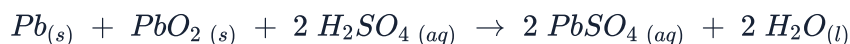
(c)



(d)



(e)



Solution

(a)



=> Oxidized substance

AgBr => Reduced substance

AgBr => Oxidizing agent



=> Reducing agent

(b)



HCHO => Oxidized substance



=> Reduced substance



=> Oxidizing agent

HCHO => Reducing agent

(c)



HCHO => Oxidized substance



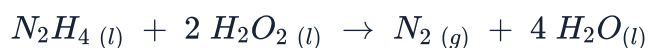
=> Reduced substance



=> Oxidizing agent

HCHO => Reducing agent

(d)



=> Oxidized substance



=> Reduced substance

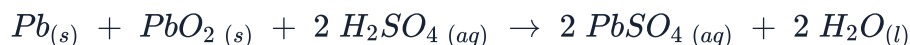


=> Oxidizing agent



=> Reducing agent

(e)



Pb=> Oxidized substance



=> Reduced substance

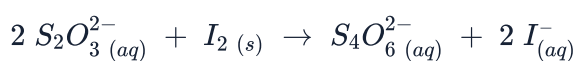


=> Oxidizing agent

Pb => Reducing agent

Question 14

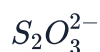
Consider the reactions :



Why does the same reductant, thiosulphate react differently with iodine and bromine ?

Solution

The average oxidation no. of S in



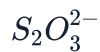
is +2.

The average oxidation no. of S in



is +2.5.

The oxidation no. of S in



is +2.

The oxidation no. of S in



is +6.

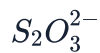
As



is a stronger oxidizing agent than



, it oxidizes S of



to a higher oxidation no. of +6 in

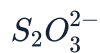


.

As



is a weaker oxidizing agent, it oxidizes S of



ion to a lower oxidation no. that is 2.5 in



ions.

Thus, thiosulphate reacts differently with



and



.

Question 15

Justify giving reactions that among halogens, fluorine is the best oxidant and among hydrohalic compounds, hydroiodic acid is the best reductant.

Solution

F_2
can oxidize



to



,



to



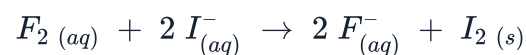
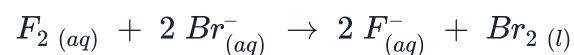
, and



to



as:



But,



,



, and



cannot oxidize



to



. The oxidizing power of halogens increases in the order as given below:



<



<



<



Therefore, fluorine is the best oxidant among halogens.



and



can reduce



to



, but



and



cannot. Hence,



and



are stronger reductants compared to



and



.



can reduce



to



, but



cannot.



Therefore, hydrochloric acid is the best reductant among hydrohalic compounds.

Hence, the reducing power of hydrohalic acids increases as given below:



<



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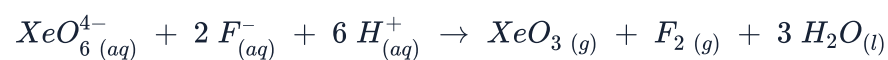


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Question 16

Why does the following reaction occur?



What conclusion about the compound

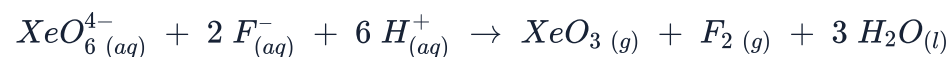


(of which



is a part) can be drawn from the reaction?

Solution



The oxidation no. of Xe reduces from +8 in



to +6 in



.

The oxidation no. of F increases from -1 in



to 0 in



.

Hence,



is reduced on the other hand



is oxidized. As



(or



) is a stronger oxidizing agent compared to

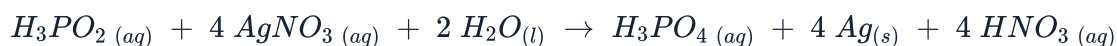


, this reaction occurs.

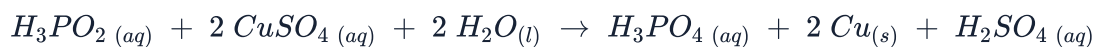
Question 17

Consider the reactions:

(a)



(b)



(c)



(d)



No change is observed

What inference do you draw about the behavior of



and



from these reactions?

Solution



and



behave as oxidizing agents in reactions (i) and (ii), respectively.

In reaction (iii),



oxidizes



to



In reaction (iv),



cannot oxidize



.

Therefore,



is a stronger oxidizing agent compared to



.

Question 18

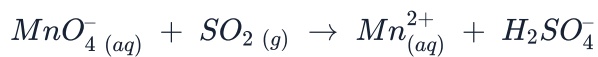
Balance the following redox reactions by ion – electron method :

(a)



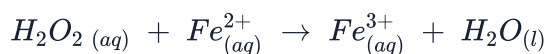
(Basic medium)

(b)



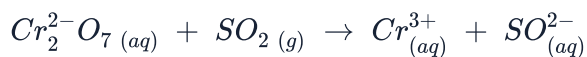
(Acidic medium)

(c)



(Acidic medium)

(d)



(Acidic medium)**Solution**

(a)



Step 1

The two half-reactions are given below:

Oxidation half-reaction:

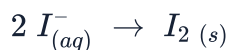


Reduction half-reaction:



Step 2

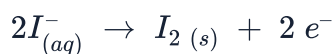
Balance I in oxidation half-reaction:



Add 2

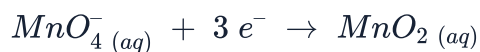


to the right-hand side of the reaction to balance the charge:



Step 3

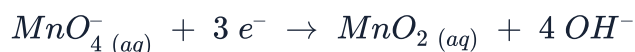
The oxidation no. of Mn has decreased from +7 to +4 in the reduction half-reaction. Therefore, 3 electrons are added to the left-hand side of the reaction.



Add 4

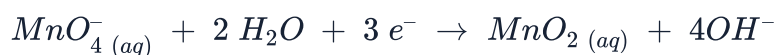


ions to the right-hand side of the reaction to balance the charge.



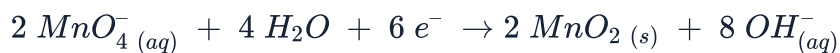
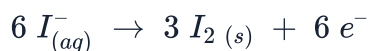
Step 4

There are 6 oxygen atoms on the right-hand side and 4 oxygen atoms on the left-hand side. Hence, 2 water molecules are added to the left-hand side.



Step 5

Equal the no. of electrons on both sides by multiplying the oxidation half-reaction by 3 and the reduction half-reaction by 2:

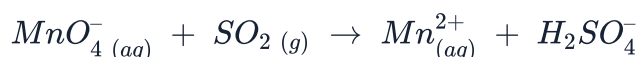


Step 6

After adding both the half-reactions, we get the balanced reaction as given below:



(b)



Step 1

Similar to (i), the oxidation half-reaction is:



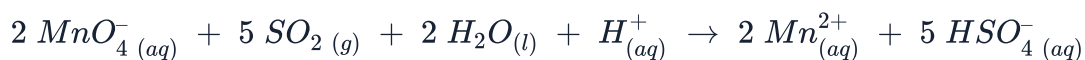
Step 2

Reduction half-reaction is:

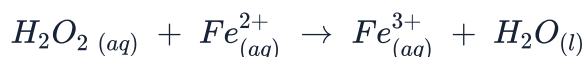


Step 3

Multiply the oxidation half-reaction with 5 and the reduction half-reaction with 2, then add them. We get the balanced reaction as given below:



(c)



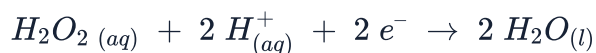
Step 1

Similar to (i), oxidation half-reaction is:



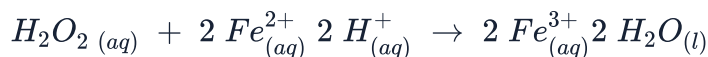
Step 2

Reduction half-reaction is:

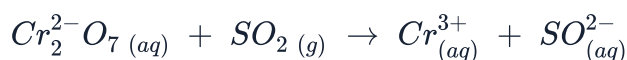


Step 3

Multiply the oxidation half-reaction with 2 then add it to the reduction half-reaction. We get the balanced reaction as given below:



(d)



Step 1

Similar to (i), oxidation half-reaction is:



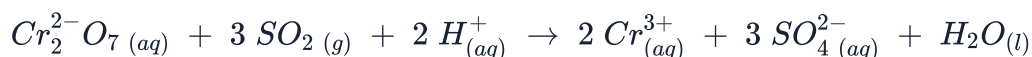
Step 2

Reduction half-reaction is:



Step 3

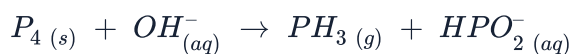
Multiply the oxidation half-reaction with 2, then add it to the reduction half-reaction. We get the balanced reaction as given below:



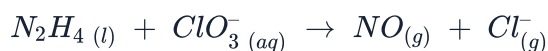
Question 19

Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidising agent and the reducing agent.

(a)



(b)



(c)

**Solution**

(a) The Oxidation no. of P reduces from 0 in



to – 3 in



The oxidation no. of P increases from 0 in



to + 2 in



. Therefore,



behaves both as a reducing agent as well as an oxidizing agent in the reaction.

Ion–electron method:

– The oxidation half-reaction:



– Balance atom P:



– Add 8 electrons to balance oxidation no.



– Add



to balance the charge:



– Add 4



to balance H and O atoms:



————(1)

– The reduction half-reaction:



– Balance atom P:



– Add 12 electrons to balance oxidation no.



– Add



to balance the charge:



– Add 12

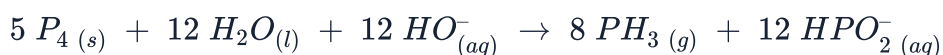


to balance H and O atoms:



———— (2)

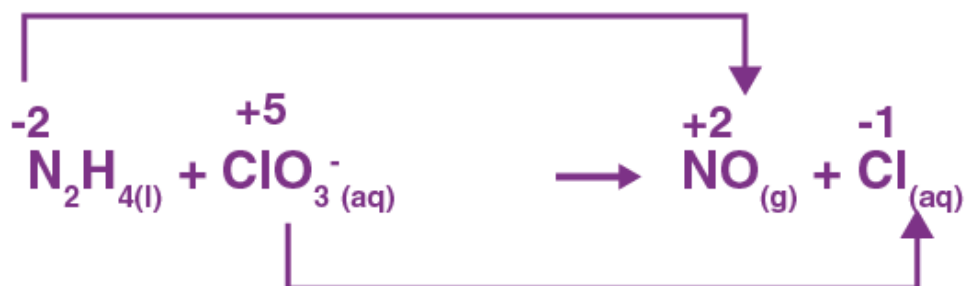
– Now, multiply the equation (1) by 3 and equation (2) by 2. Then, after adding them, we get the balanced redox reaction as given below:



(b)



O.N of M increases by 4 per atom



O.N of Cl decreases by 6 per atom

The Oxidation no. of N increases from -2 in



to -2 in NO .

The oxidation no. of Cl reduces from $+5$ in



to $+1$ in



.

Therefore,



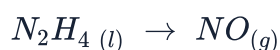
behaves as a reducing agent while



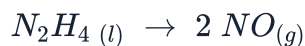
behaves as an oxidizing agent in the reaction.

Ion–electron method:

– The oxidation half-reaction:



– Balance atom N :



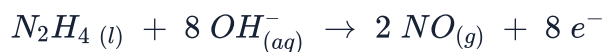
– Add 8 electrons to balance oxidation no:



– Add



to balance the charge:



– Add 6

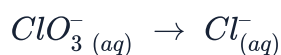


to balance O atoms:

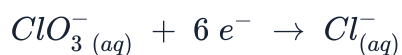


———— (1)

– The reduction half-reaction:



– Add 6 electrons to balance oxidation no.



– Add



ions to balance the charge:



– Add 3

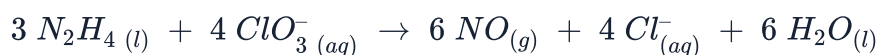


to balance O atoms:



———— (2)

Now, multiply equation (1) by 3 and equation (2) by 4. Then, after adding them, we get the balanced redox reaction as given below:



Oxidation number method:

– Reduction in the oxidation no. of N = $2 \times 4 = 8$

– Increment in the oxidation no. of Cl = $1 \times 6 = 6$

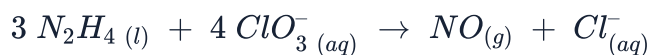
Multiply



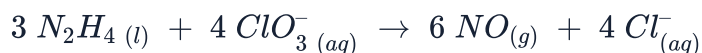
by 3 and



by 4 to balance the reduction and increment of the oxidation no. :



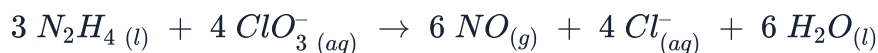
– Balance Cl and n atoms:



– Add 6

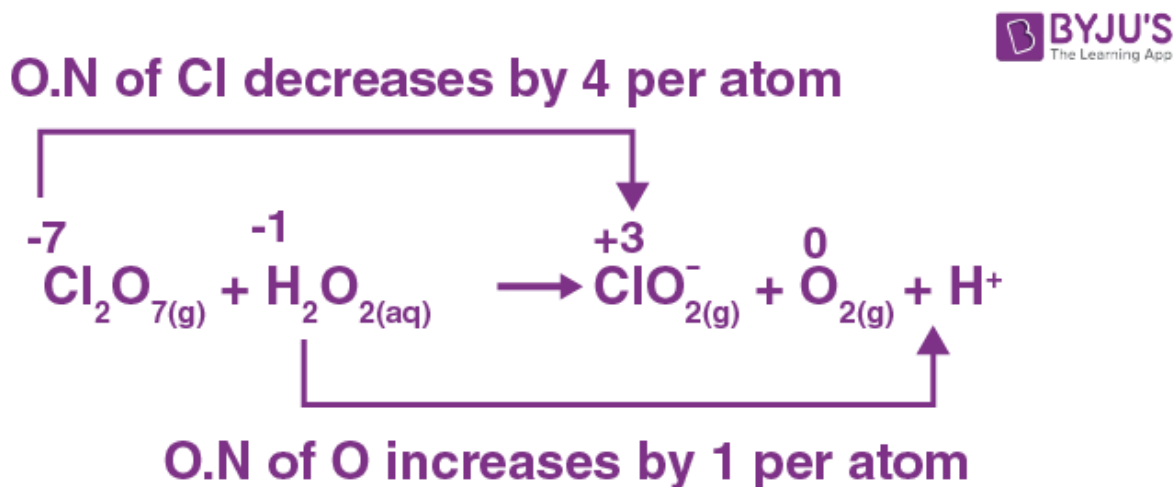


to balance O atoms:

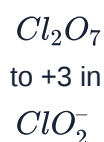


This is the required reaction equation.

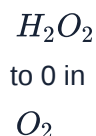
(c)



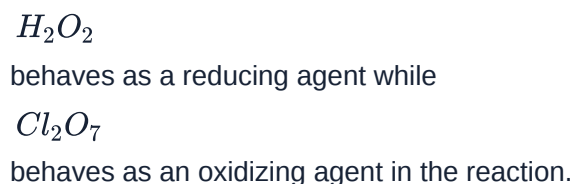
The Oxidation no. of Cl decreases from +7 in



The oxidation no. of increases from -1 in

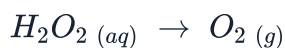


Therefore,



Ion–electron method:

- The oxidation half-reaction:



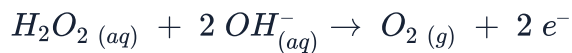
- Add 2 electrons to balance oxidation no:



– Add



to balance the charge:



– Add 2



to balance O atoms:



—— (1)

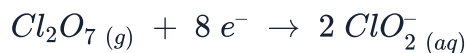
– The reduction half-reaction:



– Balance Cl atoms:



– Add 8 electrons to balance oxidation no.



– Add



ions to balance the charge:



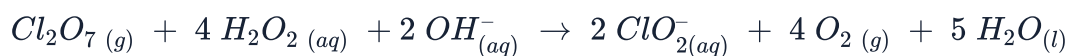
– Add 3



to balance O atoms:



Now, multiply the equation (1) by 4. Then, adding equation (1) and (2), we get the balanced redox reaction as given below:



Oxidation number method:

– Reduction in the oxidation no. of



$$= 4 \times 2 = 8$$

– Increment in the oxidation no. of



$$= 2 \times 1 = 2$$

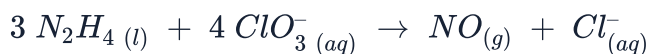
Multiply



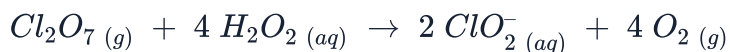
by 4 and



by 4 to balance the reduction and increment of the oxidation no. :



– Balance Cl and n atoms:



– Add 3



to balance O atoms:



– Add



and



to balance H atoms:



This is the required reaction equation.

Question 20

What sorts of informations can you draw from the following reaction ?



Solution

The oxidation no. of C in



,



and



are +3, +2 and +4 respectively.

Let the oxidation no. of C be y.



$$2(y - 3) = 0$$

Therefore, y = 3



$$y - 3 = -1$$

Therefore, y = 2



$$y - 3 - 2 = -1$$

Therefore, y = 4

The oxidation no. of C in the reaction is:

Oxidation no. of C in



is +3

Oxidation no. of C in



is +2

Oxidation no. of C in



is +4

We can see that the same compound is oxidized and reduced simultaneously in the reaction.

The reactions in which the same compound is oxidized and reduced is known as disproportionation reaction.

Then, we can say that the alkaline decomposition of cyanogens is a disproportionation reaction.

Question 21

The



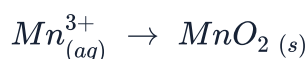
ion is unstable in solution and undergoes disproportionation to give Mn^{2+} , MnO_2 , and H^+ ion. Write a balanced ionic equation for the reaction.

Solution

The reaction is as given below:



The oxidation half-reaction:



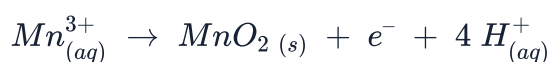
Add 1 electron to balance the oxidation no. :



Add



ions to balance the charge:



Add 2



to balance O atoms and



ions:

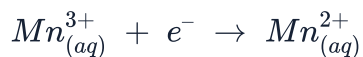


———— (1)

The reduction half-reaction:



Add 1 electron to balance oxidation no. :



———— (2)

Add equations (1) and (2) to get the balanced chemical equation:



Question 22

Consider the elements:

Cs, Ne, I and F

- (a) Identify the element that exhibits only negative oxidation.
- (b) Identify the element that exhibits only positive oxidation.
- (c) Identify the element that exhibits both negative and positive oxidation states.
- (d) Identify the element that exhibits neither negative nor positive oxidation state?

Solution

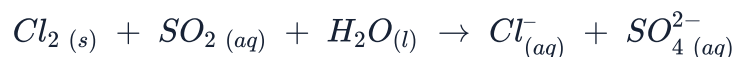
- (a) F exhibits only negative oxidation no. That is -1.
- (b) Cs exhibits only positive oxidation no. That is +1.
- (c) I exhibits both negative and positive oxidation no. That is -1, +1, +3, +5 and +7.
- (d) Ne exhibits neither negative nor positive oxidation no. That is 0.

Question 23

Chlorine is used to purify drinking water. Excess of chlorine is harmful. The excess of chlorine is removed by treating with sulphur dioxide. Present a balanced equation for this redox change taking place in water.

Solution

The redox reaction is as given below:



The oxidation half-reaction:



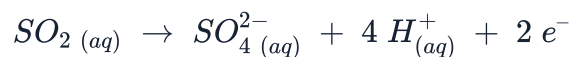
Add 2 electrons to balance the oxidation no. :



Add



ions to balance the charge:



Add 2



to balance O atoms and

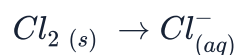


ions:



———— (1)

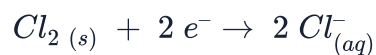
The reduction half-reaction:



Balance Cl atoms:



Add 2 electrons to balance the oxidation no. :



———— (2)

Add equations (1) and (2) to get the balanced chemical equation:



Question 24

Refer to the periodic table given in your book and now answer the following questions:

(a) Select the possible non-metals that can show disproportionation reaction?

(b) Select three metals that show disproportionation reaction?

Solution

One of the reacting elements always has an element that can exist in at least 3 oxidation numbers.

(i) The non-metals which can show disproportionation reactions are P, Cl and S.

(ii) The three metals which can show disproportionation reactions are Mn, Ga and Cu.

Question 25

In Ostwald's process for the manufacture of nitric acid, the first step involves the oxidation of ammonia gas by oxygen gas to give nitric oxide gas and steam. What is the maximum weight of nitric oxide that can be obtained starting only with 10.00 g. of ammonia and 20.00 g of oxygen?

Solution

The balanced reaction is as given below:



$$= 4 \times 17 \text{ g} = 68 \text{ g}$$



$$= 5 \times 32 \text{ g} = 160 \text{ g}$$



$$= 4 \times 30 \text{ g} = 120 \text{ g}$$



$$= 6 \times 18 \text{ g} = 108 \text{ g}$$

Thus,



(68 g) reacts with

O_2
(20 g)

Therefore, 10 g of

NH_3
reacts with
 $\frac{160 \times 10}{68}$
g = 23.53 g of

O_2

But only 20 g of

O_2
is available.

Hence,

O_2
is a limiting reagent.

Now, 160 g of

O_2
gives
 $\frac{120 \times 20}{160}$
g of N = 15 g of NO.

Therefore, max of 15 g of nitric oxide can be obtained.

Question 26

Using the standard electrode potentials given in Table 8.1, predict if the reaction between the following is feasible:

(a)

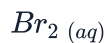
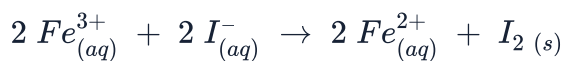


and

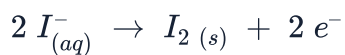


(b)



and**(c)****and****(d)****and****(e)****and****Solution****(a)****and**

Oxidation half reaction:



;

$$E^{\circ} = -0.54V$$

Reduction half reaction:



;

$$E^{\circ} = +0.77V$$



;

$$E^{\circ} = +0.23V$$

$$E^{\circ}$$

for the overall reaction is positive. Therefore, the reaction between



and



is feasible.

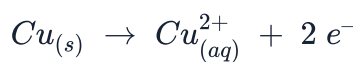
(b)



and



Oxidation half reaction:



;

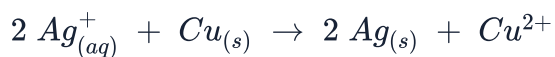
$$E^{\circ} = -0.34V$$

Reduction half reaction:



;

$$E^{\circ} = +0.80V$$



;

$$E^{\circ} = +0.46V$$

$$E^{\circ}$$

for the overall reaction is positive. Therefore, the reaction between



and



is feasible.

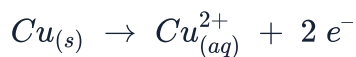
(c)



and



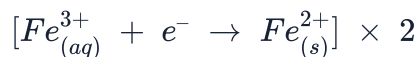
Oxidation half reaction:



;

$$E^{\circ} = -0.34V$$

Reduction half reaction:



;

$$E^{\circ} = +0.77V$$



;

$$E^{\circ} = +0.43V$$

$$E^{\circ}$$

for the overall reaction is positive. Therefore, the reaction between



and

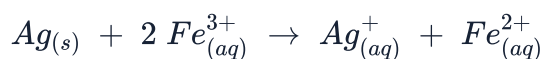


is feasible.

(d)



and



Oxidation half reaction:



;

$$E^{\circ} = -0.80V$$

Reduction half reaction:



;

$$E^{\circ} = +0.77V$$



;

$$E^{\circ} = -0.03V$$

$$E^{\circ}$$

for the overall reaction is positive. Therefore, the reaction between

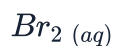


and



is feasible.

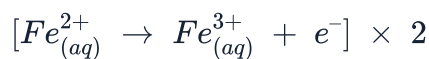
(e)



and



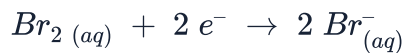
Oxidation half reaction:



;

$$E^{\circ} = -0.77V$$

Reduction half reaction:



;

$$E^{\circ} = +1.09V$$

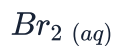


;

$$E^{\circ} = -0.32V$$

$$E^{\circ}$$

for the overall reaction is positive. Therefore, the reaction between



and



is feasible.

Question 27

Predict the products of electrolysis in each of the following:

(i) An aqueous solution of



with silver electrodes

(ii) An aqueous solution



with platinum electrodes

(iii) A dilute solution of



with platinum electrodes

(iv) An aqueous solution of



with platinum electrodes.

Solution

(i)



ionizes in aqueous solution to form



and



ions.

On electrolysis, either



ion or



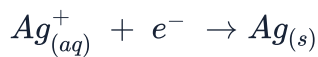
molecule can be decreased at cathode. But the reduction potential of



ions is higher than that of



.



;

$$E^{\circ} = +0.80V$$



;

$$E^{\circ} = -0.83V$$

Therefore,



ions are decreased at the cathode. Same way, Ag metal or



molecules can be oxidized at the anode. But the oxidation potential of Ag is greater than that of

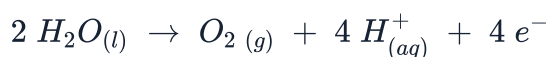


molecules.



;

$$E^{\circ} = -0.80V$$



;

$$E^{\circ} = -1.23V$$

Hence, Ag metal gets oxidized at the anode.

(ii) Pt cannot be oxidized very easily. Therefore, at the anode, oxidation of water occurs to liberate



. At the cathode,



ions are decreased and get deposited.

(iii)



ionizes in aqueous solutions to give



and



ions.



On electrolysis, either of



molecules or



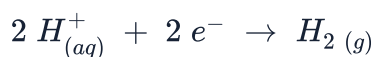
ions can get decreased at cathode. But the decreased potential of



ions is higher than that of



molecules.



;

$$E^\circ = 0.0V$$



;

$$E^\circ = -0.83V$$

Therefore, at cathode,



ions are decreased to free



gas.

On the other hand, at the anode, either of



molecules or



ions can be oxidized. But the oxidation of



involves breaking of more bonds than that of



molecules. Therefore,



ions have lower oxidation potential than



. Hence,



is oxidized at anode to free



molecules.

(iv) In aqueous solutions,



ionizes to give



and



ions as:



On electrolysis, either of



ions or



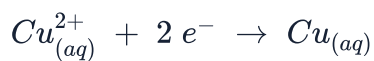
molecules can get decreased at cathode. But the decreased potential of



is more than that of



molecules.



;

$$E^{\circ} = +0.34V$$



;

$$E^{\circ} = -0.83V$$

Therefore,



ions are decreased at the cathode and get deposited. In the same way, at the anode, either of



or



is oxidized. The oxidation potential of



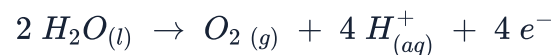
is higher than that of





;

$$E^\circ = +0.34V$$



;

$$E^\circ = -1.23V$$

But oxidation of



molecules occurs at a lower electrode potential compared to that of



ions because of over-voltage (extra voltage required to liberate gas). As a result,



ions are oxidized at the anode to liberate



gas.

Question 28

Arrange the given metals in the order in which they displace each other from the solution of their salts.

Al, Fe, Cu, Zn, Mg

Solution

A metal with stronger reducing power displaces another metal with weaker reducing power from its solution of salt.

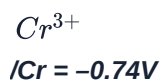
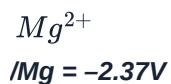
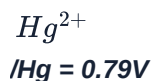
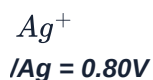
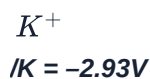
The order of the increasing reducing power of the given metals is as given below:



Therefore, Mg can displace Al from its salt solution, but Al cannot displace Mg. Thus, the order in which the given metals displace each other from the solution of their salts is as given below: $\text{Mg} > \text{Al} > \text{Zn} > \text{Fe} > \text{Cu}$

Question 29

Given the standard electrode potentials,



Arrange these metals in their increasing order of reducing power.

Solution

The reducing agent is stronger as the electrode potential decreases. Hence, the increasing order of the reducing power of the given metals is as given below:



Question 30

Depict the galvanic cell in which the reaction is:



Further show:

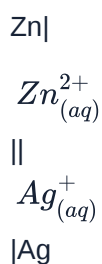
(i) which of the electrode is negatively charged?

(ii) the carriers of the current in the cell.

(iii) individual reaction at each electrode.

Solution

The galvanic cell corresponding to the given redox reaction can be shown as:



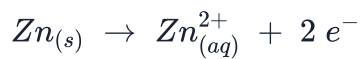
(i) Zn electrode is negatively charged because at this electrode, Zn oxidizes to



and the leaving electrons accumulate on this electrode.

(ii) The carriers of current are ions in the cell.

(iii) Reaction at Zn electrode is shown as:



Reaction at Ag electrode is shown as:

