Chapter 2: Mole Concept II

Review Questions

1. Consider the following species: Cl_2, Cl^-, ClO^-, ClO_2^-, ClO_3^-, and ClO_4^-. What is the oxidation number of Cl atom in each species?

Solution

(0, –1, +1, +3, +5, +7).

2. If you wanted to chrome plate a set of motorcycle exhaust pipes, should you be concerned that Cu^{2+} ions were also present, along with Cr^{3+} ions, in the electroplating bath?

Solution

Yes, because Cr^{3+} is more reactive than Cu.

3. MnO_4^{2-} undergoes disproportionation reaction in acidic medium but MnO_4^- does not. Give reason.

Solution

In MnO_4^-, Mn is in the highest oxidation state of +7; and hence it cannot undergo disproportionation. In contrast, the oxidation number of Mn in MnO_4^{2-} is +6. Therefore, it can increase its oxidation number to +7 or decrease its oxidation number to some lower value. Thus, MnO_4^{2-} undergoes disproportionation according to the following reaction:

$$3\overset{+6}{Mn}\overset{-2}{O_4} + 4H^+ \rightarrow 2\overset{+7}{Mn}\overset{-2}{O_4} + \overset{+4}{Mn}\overset{-2}{O_2} + 2\overset{+1}{H_2}\overset{-2}{O}$$

Here, the oxidation number of Mn increases from +6 in MnO_4^{2-} to +7 in MnO_4^- and decreases to +4 in MnO_2. Thus, MnO_4^{2-} undergoes disproportionation in acidic medium.

4. What is spectator ion? Give one example.

Solution

Spectator ions are ions that stay unaffected during a chemical reaction. They appear both as reactant and as product in an ionic equation. For example, in the following ionic equation, the sodium and nitrate ions are spectator ions.

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$$

5. Consider the equation illustrating the decomposition of hydrogen peroxide:

$$2H_2O_2(aq) + 2H_2O(l) + O_2(g)$$

What has been oxidized and what has been reduced?

Solution

In this case, H_2O_2 plays both roles. In one molecule of H_2O_2, the oxidation state of the oxygen has decreased from –1 to –2 to form H_2O, so it has been reduced. In the other molecule of H_2O_2, the oxygen has increased from –1 to 0, so it has been oxidized. One molecule of hydrogen peroxide is the oxidizing agent and the other one is the reducing agent. In such a case, we say that the H_2O_2 has been *disproportionated*.

6. Consider the following set of reactions:

$$N_2O_5 + 2NaHCO_3 \rightarrow 2NaNO_3 + 2CO_2 + H_2O$$

$$KClO_3 + 3HNO_2 \rightarrow KCl + 3HNO_3$$

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Identify which of the reactions is a redox reaction and the compounds that are oxidized and reduced.

Solution

First the oxidation numbers of all atoms must be found.

$$N_2O_5 + 3NaHCO_3 \rightarrow 2NaNO_3 + 2CO_2 + H_2O$$

Reactants:	Products:
N = +5	N = +5
O = –2	O = –2
Na = +1	
Na = +1	
H = +1	
C = +4	C = +4
O = –2	O = –2
H = +1	
O = –2	

None of the oxidation numbers change, therefore, it is not an oxidation reaction.

$KClO_3 + 3HNO_2 \rightarrow KCl + 3HNO_3$

Reactants:	Products:
K = +1	K = +1
Cl = +5	Cl = –1
O = –2	
H = +1	H = +1
N = +3	N = +5
O = –2	O = –2

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The oxidation numbers for K and Na do not change. However, the oxidation numbers for all chlorine atoms decrease. The oxidation numbers for nitrogen increase.

Therefore, $KClO_3$ is reduced and HNO_2 is oxidized.

This means $KClO_3$ is the oxidizing agent and HNO_2 is the reducing agent.

This reaction is the redox reaction. In the other reaction, the oxidation numbers of the atoms do not change.

7. Define oxidation and reduction in terms of (a) electron transfer and (b) oxidation numbers.

Solution

(a) In terms of electron transfer:

(i) **Oxidation:** The process in which an atom or an ion loses one or more electrons is known as oxidation. Thus, it is also called de-electronation.

(ii) **Reduction:** The process in which an atom or an ion gains one or more electrons is known as reduction. Thus, it is also called electronation.

(b) In terms of oxidation numbers:

(i) **Oxidation:** A chemical change in which there occurs an increase in the oxidation number of an atom or atoms is known as oxidation.

(ii) **Reduction:** A chemical change in which there occurs a decrease in the oxidation number of an

atom or atoms is known as reduction.

8. How would you identify the oxidizing agent and reducing agent in a given reaction in terms of oxidation number?

Solution

An oxidizing agent may be defined as a substance in which the oxidation number of atom (or atoms) decreases. A reducing agent may be defined as a substance in which the oxidation number of atom (or atoms) increases.

9. Show that the formation of sodium chloride from gaseous sodium and gaseous chlorine is a redox reaction.

Solution

Formation of sodium chloride from gaseous sodium and gaseous chlorine is a redox reaction. The following reaction takes place.

$$2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$$

In the oxidation half reaction, two sodium atoms lose two electrons to form two sodium cations.

$$2\text{Na} - 2e^- \rightarrow 2\text{Na}^+$$

And in the reduction half reaction, one chlorine molecule gains two electrons and forms two chloride ions.

$$\text{Cl}_2 + 2e^- \rightarrow 2\text{Cl}^-$$

10. What are the maximum and minimum oxidation numbers of N, S and Cl?

Solution

(a) The highest oxidation number of N is +5 since it has five electrons in the valence shell ($2s^2p^3$) and its minimum oxidation number is −3 since it can accept three more electrons to acquire the nearest inert gas (Ne) configuration.

(b) Similarly, the highest oxidation number of S is +6 since it has six electrons in the valence shell ($3s^23p^4$) and its minimum oxidation number is −2 since it needs only two more electrons to acquire the nearest inert gas (Ar) configuration.

(c) Likewise, the maximum oxidation number of Cl is +7 since it has seven electrons in the valence shell ($3s^23p^5$) and its minimum oxidation number is −1. So, it needs only one more electron to acquire the nearest [Ar] gas configuration.

11. Calculate the oxidation number of sulphur, chromium, and nitrogen in H_2SO_5, $Cr_2O_7^{2-}$, and NO_3^-. Suggest structure of these compounds. Account for the fallacy.

Solution

(a) The oxidation numbers of H, S, and O are $\overset{+1}{\text{H}}_2\,\overset{x}{\text{S}}\,\overset{-2}{\text{O}}_5$

Therefore,

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

$$2 + x - 10 = 0 \text{ or } x = +8$$

However, the oxidation number of S cannot be +8. S has six valence electrons. Therefore, the oxidation number of S cannot be more than +6. The structure of H_2SO_5 is shown as follows:

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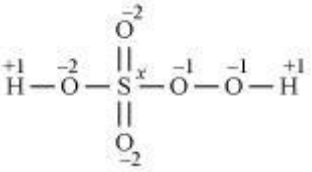


Now,

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

$$2+x-6-2=0 \text{ or } x=+6$$

Therefore, the oxidation number of S is +6.

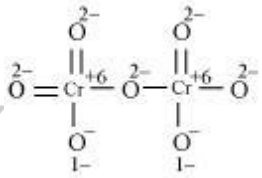
(b) The oxidation numbers of Cr and O are $\overset{x}{\mathrm{Cr}}_2\,\overset{-2}{\mathrm{O}}_7^{2-}$

Therefore,

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

$$2x-14=-2 \text{ or } x=+6$$

Here, there is no fallacy about the oxidation number of Cr in $Cr_2O_7^{2-}$. The structure of $Cr_2O_7^{2-}$ is shown as follows:



Here, each of the two Cr atoms exhibit the oxidation number of +6.

(c) The oxidation numbers of N and O are $\overset{x}{\mathrm{N}}\,\overset{2-}{\mathrm{O}}_3^-$

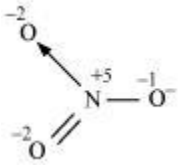
Therefore,

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

$$x-6=-1 \text{ or } x=+5$$

Here, there is no fallacy about the oxidation number of N in NO_3^-.

The structure of NO_3^- is shown as follows:



The N atom exhibits the oxidation number of +5.

12. What sorts of information can you draw from the following reaction?

$$(CN)_2(g)+2OH^-\rightarrow CN^-(aq)+CNO^-(aq)+H_2O(l)$$

Solution

The oxidation numbers of carbon in $(CN)_2$, CN^-, and CNO^- are +3, +2, and +4, respectively. These are obtained as shown below:

Let the oxidation number of C be *x*.

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$$\overset{x\ \ -3}{(CN)_2}$$

Therefore, $2(x-3)=0$ or $x=3$.

For $\overset{x\ \ -3}{(CN)^-}$, we have $x-3=-1$ or $x=2$.

For CNO^-, we have $x-3-2=-1$ or $x=4$.

The oxidation number of carbon in the various species is

$$\overset{+3}{(CN)_2}(g)+2OH^-(aq)\rightarrow \overset{+2}{C}N^-(aq)+\overset{+4}{C}NO^-(aq)+H_2O(l)$$

It can be easily observed that the same compound is being reduced and oxidized simultaneously in the given equation. Reactions in which the same compound is reduced and oxidized are known as disproportionation reactions. Thus, it can be said that the alkaline decomposition of cyanogen is an example of disproportionation reaction.

13. Why is "loss of electrons" described as oxidation?

Solution

Historically, the reaction of a substance with oxygen was termed oxidation. Now we realize that reaction with oxygen most typically means that oxygen acquires electrons from the substance with which it reacts. The oxidation of a substance is, therefore, taken to represent the loss of electrons by a substance, whether the substance has reacted with oxygen or with another oxidizing agent.

14. Laundry bleach such as Clorox is a dilute solution of sodium hypochlorite, NaOCl. Write a balanced net ionic equation for the reaction of NaOCl with $Na_2S_2O_3$. The OCl^- is reduced to chloride ion and the $S_2O_3^{2-}$ is oxidized to sulphate ion.

Solution

$(OCl^- + 2H^+ + 2e^- \rightarrow Cl^- + H_2O) \times 4$

$S_2O_3^{2-} + 5H_2O \rightarrow 2SO_4^{2-} + 10H^+ + 8e^-$

$4OCl^- + S_2O_3^{2-} + 5H_2O + 8H^+ \rightarrow 4Cl^- + 2SO_4^{2-} + 10H^+ + 4H_2O$

which simplifies to

$4OCl^- + S_2O_3^{2-} + H_2O \rightarrow 4Cl^- + 2SO_4^{2-} + 2H^+$

15. Use Table 10.2 to predict the outcome of the following reactions. If no reaction occurs, write N.R. If a reaction occurs, write a balanced chemical equation for it.

(a) $Fe + Mg^{2+} \rightarrow$ (b) $Cr + Pb^{2+} \rightarrow$

(c) $Ag^+ + Fe \rightarrow$ (d) $Ag + Au^{3+} \rightarrow$

Solution

In each case, the reaction should proceed to give the less reactive of the two metals, together with the ion of the more reactive of the two metals. The reactivity is taken from the reactivity series in Table 10.2.

(a) N.R.

(b) $2Cr(s) + 3Pb^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Pb(s)$

(c) $2Ag^+(aq) + Fe(s) \rightarrow 2Ag(s) + Fe^{2+}(aq)$

(d) $3Ag(s) + Au^{3+}(aq) \rightarrow Au(s) + 3Ag^+(aq)$

16. It is observed that when magnesium metal is dipped into a solution of nickel(II) chloride, some of the magnesium dissolves and nickel metal is deposited on the surface of the magnesium. Which of the following reactions will occur spontaneously? Explain the reason for your answer.

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(a) $2Mo^{3+} + 3Mg \rightarrow 3Mg^{2+} + 2Mo$
(b) $2Mo + 3Mg^{2+} \rightarrow 2Mo^{3+} + 3Mg$

Solution

The observation shows that Mg is more active than Ni, however, the information is not sufficient to determine which is easier to oxidize, Mg or Mo. Therefore, it cannot be determined which reaction will occur spontaneously.

17. The following chemical reactions are *observed to occur* in aqueous solution.

$$2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$$
$$2Al + 3Fe^{2+} \rightarrow 3Fe + 2Al^{3+}$$
$$Pb^{2+} + Fe \rightarrow Pb + Fe^{2+}$$
$$Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$$
$$2Al + 3Pb^{2+} \rightarrow 3Pb + 2Al^{3+}$$
$$Pb + Cu^{2+} \rightarrow Pb^{2+} + Cu$$

Arrange the metals Al, Pb, Fe, and Cu in order of increasing ease of oxidation.

Solution

Reaction 1 demonstrates that Al is more readily oxidized than Cu. Reaction 2 demonstrates that Al is more readily oxidized than Fe. Reaction 3 demonstrates that Fe is more readily oxidized than Pb. Reaction 4 demonstrates that Fe is more readily oxidized than Cu. Reaction 5 demonstrates that Al is more readily oxidized than Pb. The last reaction demonstrates that Pb is more readily oxidized than Cu. Altogether, the above facts constitute the following trend of increasing ease of oxidation:

$$Cu < Pb < Fe < Al$$

18. A solution contains $Ce(SO_4)_3^{2-}$ at a concentration of 0.0150 M. It was found that in a titration, 25.00 mL of this solution reacted completely with 23.44 mL of 0.032 M $FeSO_4$ solution. The reaction gave Fe^{3+} as a product in the solution. In this reaction, what is the final oxidation state of the Ce?

Solution

The oxidation state of cerium in the reactant ion is +4. The number of moles of this ion in the reactant solution is

$$0.0150\ M \times 0.02500\ L = 3.75 \times 10^{-4}\ mol\ Ce^{4+}$$

The number of moles of electrons that come from the Fe^{2+} reducing agent is

$$0.0320\ M \times 0.02344\ L = 7.50 \times 10^{-4}\ mol\ e^{-}$$

The ratio of moles of electrons to moles of Ce^{4+} reactant is, therefore, 2:1, and we conclude that the product is Ce^{2+}.

19. Organic chemists often think about oxidation in terms of the loss of a pair of hydrogen atoms. Ethyl alcohol, CH_3CH_2OH, is oxidized to acetaldehyde CH_3CHO, for example, by removing a pair of hydrogen atoms. Show that this reaction obeys the general rule that oxidation is any process in which the oxidation number of an atom increases.

Solution

The reaction is as follows:

$$CH_3CH_2OH \rightarrow CH_3CHO + 2e^{-} + 2H^{+}$$

The net oxidation state of the carbon next to the oxygen atom in ethanol is assigned a value of –1. The carbon atom next to the oxygen atom in acetaldehyde is assigned a net oxidation state of +1. The

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oxidation state of carbon increases from −1 to +1. This is equivalent to the loss of two electrons or the removal of two hydrogen atoms. Ethanol is oxidized to acetaldehyde.

20. Differentiate between valency and oxidation number.

Solution

Valency is the combining capacity of an element while oxidation number is the charge on an atom when present in combined state.

21. Give proper reasoning for the following:

(a) An acidified potassium dichromate paper on being exposed to sulphur dioxide turns green.

(b) Iodide salts liberate iodine on addition of Cl_2 water.

(c) $HgCl_2$ and $SnCl_2$ cannot exist as such if present together in an aqueous solution.

Solution

(a) $K_2Cr_2O_7$ is reduced to green $Cr_2(SO_4)_3$ according to the following equation

$$K_2Cr_2O_7 + 3SO_2 + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + H_2O$$

(b) Iodide salts are oxidized by Cl_2 water

$$2KI + Cl_2 \rightarrow 2KCl + I_2$$

(c) Both undergo redox reaction:

$$2HgCl_2 + SnCl_2 \rightarrow SnCl_4 + \underset{\text{White}}{Hg_2Cl_2} \xrightarrow{SnCl_2} \underset{\text{Gray}}{Hg_2} + SnCl_4$$

22. The color of KI solution containing starch turns blue when it is shaken with cold Cl_2 water. Explain why?

Solution

Cl_2 is a stronger oxidizing agent than I_2, therefore, when Cl_2 is passed through KI solution, Cl_2 gets reduced to Cl^- ions while I^- ions get oxidized to violet colored iodine according to the equation

$$Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(s)$$

The iodine thus produced forms a blue colored inclusion complex with starch, and hence the solution turns blue.

23. 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.

Solution

In the manufacture of benzoic acid from toluene, alcoholic potassium permanganate is used as an oxidant because of the following reasons:

(i) In a neutral medium, OH^- ions are produced in the reaction itself. As a result, the cost of adding an acid or a base can be reduced.

(ii) $KMnO_4$ and alcohol are homogeneous to each other since both are polar. Toluene and alcohol are also homogeneous to each other because both are organic compounds. Reactions can proceed at a faster rate in a homogeneous medium than in a heterogeneous medium. Hence, in alcohol, $KMnO_4$ and toluene can react at a faster rate.

The balanced redox equation for the reaction in a neutral medium is given as follows:

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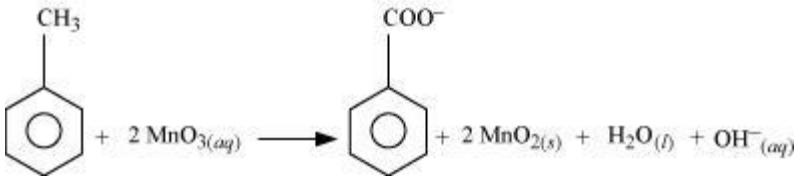

24. Chlorine is used to purify drinking water. Excess of chlorine is harmful. The excess of chlorine is removed by treating with sulphur dioxide. Present a balanced equation for this redox change taking place in water.

Solution

The given redox reaction can be represented as

$$Cl_2(s) + SO_2(aq) + H_2O(l) \rightarrow Cl^-(aq) + SO_4^{2-}(aq)$$

The oxidation half-cell reaction is

$$\overset{+4}{S}O_2(aq) \rightarrow \overset{+6}{S}O_4^{2-}(aq)$$

The oxidation number is balanced by adding two electrons as

$$SO_2(aq) \rightarrow SO_4^{2-}(aq) + 2e^-$$

The charge is balanced by adding $4H^+$ ions as

$$SO_2(aq) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$$

The O atoms and H^+ ions are balanced by adding $2H_2O$ molecules as

$$SO_2(aq) + 2H_2O(l) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \quad (1)$$

The reduction half-cell reaction is

$$Cl_2(s) \rightarrow Cl^-(aq)$$

The chlorine atoms are balanced as

$$\overset{0}{Cl_2}(s) \rightarrow \overset{-1}{Cl^-}(aq)$$

The oxidation number is balanced by adding electrons

$$Cl_2(s) + 2e^- \rightarrow 2Cl^-(aq) \quad (2)$$

The balanced chemical equation can be obtained by adding Eqs. (1) and (2) as

$$Cl_2(s) + SO_2(aq) + 2H_2O(l) \rightarrow 2Cl^-(aq) + SO_4^{2-}(aq) + 4H^+(aq)$$

25. Give the balanced equations for the reactions that form the basis for the following redox titrations:

(a) Titration of potassium permanganate for estimation of ferrous ions.

(b) Titration of iodine against sodium thiosulphate (iodimetric titration).

(c) Iodometric titration for estimation of cupric ions.

Solution

(a) $5Fe^{2+}(aq) + MnO_4^-(aq) + 8H^+(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$

(b) $I_2^-(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

(c) $2Cu^{2+}(aq) + 4I^-(aq) \rightarrow Cu_2I_2(s) + I_2(s)$

26. Define equivalent weight of an oxidizing/reducing agent.

Solution

The equivalent weight of an oxidizing/reducing agent is equal to its molecular weight divided by the

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number of electrons gained/lost as represented in the balanced chemical equation.

27. Given the following reaction: $S_2O_3^{2-}(aq) \rightarrow SO_4^{2-}(aq)$, how many H^+ ions, and on which side of the $\rightarrow$ are required to balance the half reaction?

Solution

The balanced chemical equation is

$$S_2O_3^{2-}(aq) + 5H_2O \rightarrow 2SO_4^{2-} + 10H^+ + 8e^-$$

$10H^+$ ions are required on the product side of the reaction.

28. Identify the oxidants and reductants in the following reactions.

(a) $CH_4(g) + 4Cl_2(g) \rightarrow CCl_4(g) + 4HCl$

(b) $2H^+(aq) + MnO_2(s) + C_2H_2O_4 \rightarrow Mn^{2+}(aq) + 2CO_2(g) + 2H_2O(l)$

(c) $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$.

(d) $Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$

Solution

(a) $\overset{-4}{C}H_4(g) + 4\overset{0}{Cl}_2(g) \rightarrow \overset{+4}{C}Cl_4(g) + 4\overset{-1}{H}Cl$

Oxidant: $Cl_2(g)$

Reductant: $CH_4(g)$

(b) $2H^+(aq) + \overset{+4}{Mn}O_2(s) + \overset{+3}{C}_2H_2O_4 \rightarrow Mn^{2+} + 2\overset{+4}{C}O_2(g) + 2H_2O(l)$

Oxidant: $MnO_2(s)$

Reductant: $C_2H_2O_4$

(c) $\overset{0}{I}_2(aq) + 2\overset{+2}{S}_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + \overset{+2.5}{S}_4O_6^{2-}(aq)$

Oxidant: $I_2(aq)$

Reductant: $S_2O_3^{2-}(aq)$

(d) $\overset{0}{Cl}_2(g) + 2\overset{-1}{Br}^-(aq) \rightarrow 2\overset{-1}{Cl}^-(aq) + \overset{0}{Br}_2(aq)$

Oxidant: $Cl_2(g)$

Reductant: $Br^-(aq)$

29. What are half reactions? Explain with examples.

Solution

Half reactions can be oxidation half reactions where only oxidation is involved and reduction half reaction involving reduction only.

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

Reduction half reaction: $Cu^{2+} + 2e^- \rightarrow Cu$

30. H_2O_2 acts as reductant as well as oxidant. Explain.

Solution

Oxidizing action of H_2O_2: $4H_2O_2 + PbS \rightarrow PbSO_4 + 4H_2O$

Reducing action of H_2O_2: $PbO_2 + H_2O_2 \rightarrow PbO + H_2O + O_2$

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31. Balance the following equations in an acidic medium by both oxidation number and ion-electron methods and identify the oxidants and reductants:

(a) $Bi(s) + NO_3^-(aq) \rightarrow Bi^{3+}(aq) + NO_2(g)$

(b) $H_2S(g) + Fe^{3+}(aq) \rightarrow Fe^{2+} + S(s) + H^+(aq)$

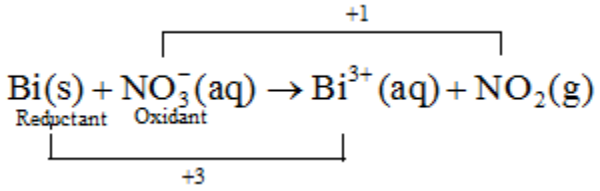
(c) $S(s) + HNO_3(aq) \rightarrow SO_2(g) + NO_2(g) + H_2O(l)$

(d) $As(s) + NO_3^-(aq) + H^+(aq) \rightarrow AsO_4^{3-}(aq) + NO_2(g) + H_2O(l)$

Solution

(a) $Bi(s) + NO_3^-(aq) \rightarrow Bi^{3+}(aq) + NO_2(g)$

By oxidation number method:



Equalize increase/decrease of oxidation number by multiplying NO_3^- by 3 and Bi by 1:

$$Bi + 3NO_3^- \rightarrow Bi^{3+} + 3NO_2$$

Balance O atoms:

$$Bi + 3NO_3^- \rightarrow Bi^{3+} + 3NO_2 + 3H_2O$$

Balance H atoms in acidic medium:

$$Bi + 3NO_3^- + 6H^+ \rightarrow Bi^{3+} + 3NO_2 + 3H_2O$$

By ion-electron method:

$$Bi \rightarrow Bi^{3+} + 3e^- \text{ Oxidation half}$$

$$NO_3^- + e^- \rightarrow NO_2 \text{ Reduction half}$$

Then balance O atoms:

$$NO_3^- + e^- + 2H^+ \rightarrow NO_2 + H_2O \text{ Reduction half}$$

$$Bi \rightarrow Bi^{3+} + 3e^- \text{ Oxidation half}$$

Equalize the electrons by multiplying the reduction half with 5 and the two half reactions:

$$NO_3^- + e^- + 2H^+ \rightarrow NO_2 + H_2O \times 3$$

$$Bi \rightarrow Bi^{3+} + 3e^-$$

$$Bi + 3NO_3^- + 6H^+ \rightarrow Bi^{3+} + 3NO_2 + 3H_2O$$

(b) $H_2S(g) + Fe^{3+}(aq) \rightarrow Fe^{2+} + S(s) + H^+(aq)$

By oxidation number method:

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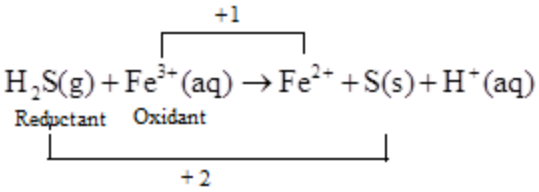


To equalize oxidation number multiple Fe^{3+} by 2 and H_2S by 1

$$\overset{+1}{H_2}\overset{-2}{S}(g) + 2Fe^{3+}(aq) \rightarrow 2Fe^{2+} + \overset{0}{S}(s) + 2H^+(aq)$$

By ion-electron method:

$$Fe^{3+} + e^- \rightarrow Fe^{2+} \quad (1)$$

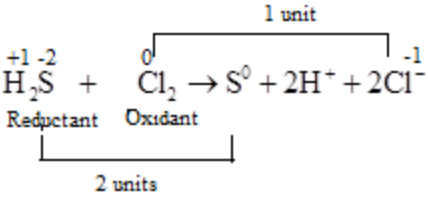
$$H_2S \rightarrow S + 2H^+ + 2e^- \quad (2)$$

To equalize electrons multiply Eq. (1) by (2) and (2) by (1)

$$Fe^{3+} + e^- \rightarrow Fe^{2+}]\times 2$$

$$\frac{H_2S \rightarrow S + 2H^+ + 2e^-}{2Fe^{3+} + H_2S \rightarrow S(s) + Cl^-(aq)}$$

By oxidation number method:



Balance the reaction:

$$H_2S + Cl_2 \rightarrow S^0 + 2H^+ + 2Cl^-$$

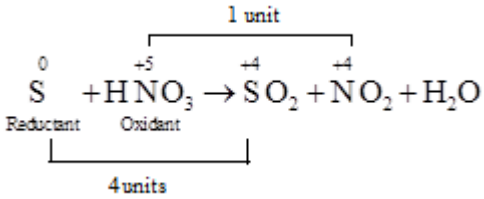
By ion-electron method:

$$H_2S \rightarrow S + 2e^- + 2H^+$$

$$\frac{Cl_2 + 2e^- \rightarrow 2Cl^-}{H_2S + Cl_2 \rightarrow S + 2Cl^- + 2H^+}$$

(c) $S + HNO_3 \rightarrow SO_2 + NO_2 + H_2O$

By oxidation number method:



Equalize increase/decrease of oxidation number by multiplying HNO_3 by 4 and S by 1

$$4HNO_3 \rightarrow 4NO_2 + 4H_2O$$

$$S \rightarrow SO_2$$

Balance O atoms:

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$$4HNO_3 \rightarrow 4NO_2 + 4H_2O$$
$$S + 2H_2O \rightarrow SO_2$$

Balance H atoms in acidic medium:

$$4HNO_3 + 4H^+ \rightarrow 4NO_2 + 4H_2O$$
$$\underline{S + 2H_2O \rightarrow SO_2 + 4H^+}$$
$$4HNO_3 + S \rightarrow 4NO_2 + SO_2 + 2H_2O$$

By ion-electron method:

$$S \rightarrow \overset{+4}{S}O_2 + 4e^- \quad \text{Oxidation half}$$
$$H\overset{+5}{N}O_3 + e^- \rightarrow \overset{+4}{N}O_2 + H_2O \quad \text{Reduction half}$$

Then balance O atoms and H atoms:

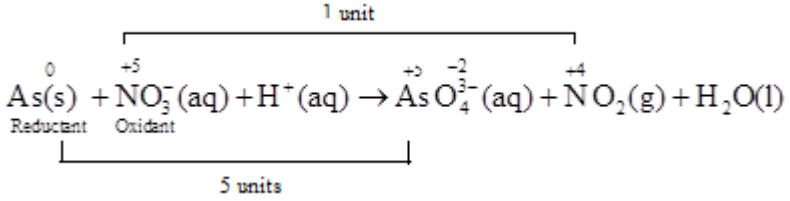
$$S + 2H_2O \rightarrow SO_2 + 4e^- + 4H^+$$
$$H^+ + HNO_3 + e^- \rightarrow NO_2 + H_2O$$

Equalize the electrons:

$$S + 2H_2O \rightarrow SO_2 + 4e^- + 4H^+$$
$$\underline{H^+ + HNO_3 + e^- \rightarrow NO_2 + H_2O \times 4}$$
$$S + 4HNO_3 \rightarrow 4NO_2 + SO_2 + 2H_2O$$

(d) $As(s) + NO_3^-(aq) + H^+(aq) \rightarrow AsO_4^{3-}(aq) + NO_2(g) + H_2O(l)$

By oxidation number method:



Equalize increase/decrease of O.N. by multiplying NO_3^- by 5 and As by 1

$$As(s) + 5NO_3^-(aq) + H^+(aq) \rightarrow AsO_4^{3-}(aq) + 5NO_2(g) + H_2O(l)$$

Balance O atom and H atom in acidic medium

$$As(s) + 5NO_3^-(aq) + 2H^+(aq) \rightarrow AsO_4^{3-}(aq) + 5NO_2(g) + H_2O(l)$$

By ion-electron method:

$$As \rightarrow AsO_4^{3-} + 5e^- \quad \text{Oxidation half reaction}$$
$$NO_3^- + e^- \rightarrow NO_2 \quad \text{Reduction half}$$

Then balance O and H atoms:

$$As + 4H_2O \rightarrow AsO_4^{3-} + 5e^- + 8H^+$$
$$NO_3^- + e^- + 2H^+ \rightarrow NO_2 + H_2O$$

Equalize the electrons by multiplying the reduction half by 5 and oxidation half by 1

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$$As + 4H_2O \rightarrow AsO_4^{3-} + 5e^- + 8H^+$$
$$\underline{NO_3^- + e^- + 2H^+ \rightarrow NO_2 + H_2O \times 5}$$
$$As + 5NO_3^- + 2H^+ \rightarrow AsO_4^{3-} + 5NO_2 + H_2O$$

32. Identify the substances oxidized and reduced and the oxidizing and reducing agents in the reaction of (a) sodium with molecular oxygen O_2; (b) aluminium with chlorine.

Solution

(a) Sodium reacts with molecular oxygen to form sodium peroxide Na_2O_2, which contains the peroxide ion O_2^{2-}. The reaction involved is

$$Na(s) + O_2(g) \rightarrow Na_2O_2(s)$$

Oxygen is reduced since it gains electrons, and so acts as an oxidizing agent, whereas sodium is oxidized since it loses electrons and so acts as a reducing agent.

(b) Aluminium reacts with chlorine to form aluminium chloride. The reaction is

$$2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(aq)$$

Aluminium is oxidized and is, therefore, the reducing agent. Chlorine is reduced and is, therefore, the oxidizing agent.

33. Balance equations for these reactions occurring in a basic solution.

(a) $CrO_4^{2-} + S^{2-} \rightarrow S + CrO_2^-$

(b) $MnO_4^- + C_2O_4^{2-} \rightarrow CO_2 + MnO_2$

(c) $ClO_3^- + N_2H_4 \rightarrow NO + Cl^-$

Solution

For redox reactions in basic solution, we proceed to balance the half reactions as if they were in acid solution, and then add enough OH^- to each side of the resulting equation in order to neutralize (titrate) all of the H^+. This gives a corresponding amount of water ($H^+ + OH^- \rightarrow H_2O$) on one side of the equation, and an excess of OH^- on the other side of the equation, as befits a reaction in basic solution.

(a) $(CrO_4^{2-} + 4H^+ + 3e^- \rightarrow CrO_2^- + 2H_2O) \times 2$

$(S^{2-} \rightarrow S + 2e^-) \times 3$

$2CrO_4^{2-} + 3S^{2-} + 8H^+ \rightarrow 2CrO_2^- + 3S + 4H_2O$

Adding $8OH^-$ to both sides of the above equation we obtain

$2CrO_4^{2-} + 3S^{2-} + 8H_2O \rightarrow 2CrO_2^- + 8OH^- + 3S + 4H_2O$

which simplifies to

$2CrO_4^{2-} + 3S^{2-} + 4H_2O \rightarrow 2CrO_2^- + 3S + 8OH^-$

(b) $(C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-) \times 3$

$(MnO_4^- + 4H^+ + 3e^- \rightarrow\rightarrow MnO_2 + 2H_2O) \times 2$

$3C_2O_4^{2-} + 2MnO_4^- + 8H^+ \rightarrow 6CO_2 + 2MnO_2 + 4H_2O$

Adding $8OH^-$ to both sides of the above equation we get

$3C_2O_4^{2-} + 2MnO_4^- + 8H_2O \rightarrow 6CO_2 + 2MnO_2 + 4H_2O + 8OH^-$

which simplifies to give

$3C_2O_4^{2-} + 2MnO_4^- + 4H_2O \rightarrow 6CO_2 + 2MnO_2 + 8OH^-$

(c) $(ClO_3^- + 6H^+ + 6e^- \rightarrow Cl^- + 3H_2O) \times 4$

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$(N_2H_4 + 2H_2O \rightarrow 2NO + 8H^+ + 8e^-) \times 3$

$4ClO_3^- + 3N_2H_4 + 24H^+ + 6H_2O \rightarrow 4Cl^- + 6NO + 12H_2O + 24H^+$

which needs no OH^-, because it simplifies directly to

$4ClO_3^- + 3N_2H_4 \rightarrow 4Cl^- + 6NO + 6H_2O$

34. What is the oxidation state of Cl in (A) $HClO_4$, (B) $Ba(Cl_3)_2$, (C) CrO_2Cl_2, and (D) Cl_2O_7?

Solution

(A) +7 (B) +5 (C) –1 (D) +7

35. Indicate the species that are oxidized and reduced in the following reactions.

(A) $CN^- + OCl^- \rightarrow N_2 + HCO_3^- + Cl^-$

(B) $C_6H_5OH + O_3 \rightarrow CO_2 + H_2O + O_2$

(C) $I^- + O_3 \rightarrow O_2 + I_2$

(D) $S_2O_3^{2-} + Cl_2 \rightarrow HSO_4^- + Cl^-$

(E) $C + ZnO \rightarrow Zn + CO$

Solution

	(A)	(B)	(C)	(D)	(E)
Oxidized	CN^-	C_2H_5OH	I^-	$S_2O_3^{2-}$	C
Reduced	OCl^-	O_3	O_3	Cl_2	ZnO

36**.** What is implied by the statement "oleum is labeled 109% H_2SO_4"?

Solution

Oleum is also called pyrosulphuric acid, $H_2S_2O_7$. It is obtained by dissolving SO_3 in concentrated $H_2S_2O_4$

$$H_2SO_4 + SO_3 \rightarrow H_2S_2O_7$$

Therefore, 109% H_2SO_4 means strength of oleum in terms of sulphuric acid potential and 9 g of free SO_3 in 100 g oleum.

Numerical Problems

1. 0.5 g of sample containing MnO_2 is treated with HCl, liberating Cl_2. The Cl_2 is passed into a solution of KI and 30.0 cm^3 of 0.1 M $Na_2S_2O_3$ which is required to titrate the liberated iodine. Calculate the percentage of MnO_2 in sample (atomic weight of Mn = 55 u).

Solution

The reaction involved is

$$MnO_2 \xrightarrow{HCl} Cl_2 \xrightarrow{KI} I_2 \xrightarrow{Na_2S_2O_3} NaI + Na_2S_4O_6$$

Redox changes involved are

$$2e^- + I_2^0 \rightarrow 2I^-$$

$$2S_2^{2+} \rightarrow S_4^{(5/2)^+} + 2e^-$$

$$Mn^{4+} + 2e^- \rightarrow Mn^{2+}$$

The reactions suggest that Milliequiv. of MnO_2 = Milliequiv. of Cl_2 formed = Milliequiv. of I_2 liberated = Milliequiv. of $Na_2S_2O_3$ used

$$\text{Therefore, } \frac{w}{M/2} \times 1000 = 0.1 \times 1 \times 30$$

(as $N_{Na_2S_2O_3} = M_{Na_2S_2O_3}$ since valence factor = 1, see redox changes for $Na_2S_2O_3$)

or $w_{MnO_2} = \frac{0.1 \times 1 \times 30 \times M}{2000} = \frac{0.1 \times 1 \times 30 \times 87}{2000} = 0.1305\,g$

(as $M_{MnO_2} = 87\,g\,mol^{-1}$)

Hence, percentage purity of MnO_2 is

$$\text{Percentage of } MnO_2 = \frac{0.1305}{0.5} \times 100 = 26.1\%$$

2. The neutralization of a solution of 1.2 g of mixture of $H_2C_2O_4 \cdot 2H_2O$ and $KHC_2O_4 \cdot H_2O$ and some impurities which are neutral, consumed 40.0 mL 0.25 N NaOH. On the other hand, on titration with $KMnO_4$, for a 0.4 g of mixture required 40.0 mL of 0.125 N $KMnO_4$. Find the percentage composition of mixture.

Solution

Note that the mass of two samples is different in two titrations.

Let x = millimoles of $H_2C_2O_4 \cdot 2H_2O$ and y = millimoles of $KHC_2O_4 \cdot H_2O$ in 0.4 g sample

So, the millimoles of oxalic acid in 1.2 g sample = $3x$

And millimoles of bioxalate in 1.2 g sample = $3y$

(a) **Neutralization with NaOH:** In this case, milliequiv. of $H_2C_2O_4 \cdot 2H_2O$ + milliequiv. of $KHC_2O_4 \cdot H_2O$ = milliequiv. of NaOH, or

$2 \times 3x + 1 \times 3y = 0.25 \times 40$ (acid has 2 H^+ ions while bioxalate has only 1 H^+) $= 2x + y = 10/3$

(b) **Redox reaction with $KMnO_4$:** In this case, milliequiv. of $H_2C_2O_4 \cdot 2H_2O$ + milliequiv. of $KHC_2O_4 \cdot H_2O$ = milliequiv. of $KMnO_4$, or

$$2 \times x + 2 \times y = 0.125 \times 40$$

From the reaction, $HC_2O_4^- \rightarrow 2CO_2 + H^+ + 2e^-$ we have $x = 2$.

And from the reaction $C_2O_4^{2-} \rightarrow 2CO_2 + H^+ + 2e^-$, we have $x + y = 2.5$

Solving for x and y, we get $x = 5/6$ and $y = 5/3$.

Using

$$\frac{g_{H_2C_2O_4 \cdot 2H_2O}}{M_{H_2C_2O_4 \cdot 2H_2O}} \times 1000 = \frac{5}{6} \quad \text{and} \quad \frac{g_{KHC_2O_4 \cdot 2H_2O}}{M_{KHC_2O_4 \cdot 2H_2O}} \times 1000 = \frac{5}{3}$$

we get

$$g_{H_2C_2O_4 \cdot 2H_2O} = 0.105\ g \quad \text{and} \quad g_{KHC_2O_4 \cdot 2H_2O} = 0.243\ g$$

Therefore,

$$\text{Percentage of } H_2C_2O_4 \cdot 2H_2O = \frac{0.105}{0.4} \times 100 = 26.25\% \text{ and percentage of } KHC_2O_4 \cdot H_2O = \frac{0.243}{0.4} \times 100 = 60.75\%$$

3. Iodate ion reacts with sulphite ion to give sulphate ion and iodide ion.

(a) Write a balanced net ionic equation for the reaction.

(b) How many grams of sodium sulphite are needed to react with 5.00 g of sodium iodate?

Solution

(a) $IO_3^- + 6H^+ + 6e^- \rightarrow I^- + 3H_2O$

$[SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2e^-] \times 3$

$IO_3^- + 3SO_3^{2-} + 6H^+ + 3H_2O \rightarrow I^- + 3SO_4^{2-} + 3H_2O + 6H^+$

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which simplifies to

$IO_3^- + 3SO_3^{2-} \rightarrow I^- + 3SO_4^{2-}$

(b) g Na_2SO_3 = (5.00 g $NaIO_3$)$\left(\frac{1\,\text{mol NaIO}_3}{197.9\,\text{g NaIO}_3}\right)\left(\frac{3\,\text{mol Na}_2\text{SO}_3}{1\,\text{mol NaIO}_3}\right)\left(\frac{126.0\,\text{g Na}_2\text{SO}_3}{1\,\text{mol Na}_2\text{SO}_3}\right)$ = 9.55 g Na_2SO_3

4. An organic compound contains carbon, hydrogen, and sulphur. A sample of it with a mass of 1.045 g was burned in oxygen to give gaseous CO_2, H_2O and SO_2. These gases were passed through 500.0 mL of an acidified 0.0200 M $KMnO_4$ solution, which caused the SO_2 to be oxidized to SO_4^{2-}. Only part of the available $KMnO_4$ was reduced to Mn^{2+}. Next, 50.00 mL of 0.0300 M $SnCl_2$ was added to a 50.00 mL portion of this solution, which still contained unreduced $KMnO_4$. There was more than enough added $SnCl_2$ to cause all of the remaining MnO_4^- in the 50 mL portion to be reduced to Mn^{2+}. The excess Sn^{2+} that still remained after the reaction was then titrated with 0.0100 M $KMnO_4$, requiring 27.28 mL of the $KMnO_4$ solution to reach the end point. What was the percentage of sulphur in the original sample of the organic compound that had been burned?

Solution

It is first necessary to write a balanced equation for the reaction of MnO_4^- with Sn^{2+}.

$Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

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

$5Sn^{2+} + 2\,MnO_4^- + 16H^+ \rightarrow 5Sn^{4+} + 2Mn^{2+} + 8H_2O$

Then, we calculate the original moles of Sn^{2+} in the 50.0 mL of 0.0300 M $SnCl_2$ solution.

Mol Sn^{2+} = (50.0 mL Sn^{2+})$\left(\frac{0.0300\ \text{mol Sn}^{2+}}{1000\ \text{mL}}\right)$ = 0.0015 mol Sn^{2+}

The moles of Sn^{2+} that were titrated are calculated by multiplying (remembering to include stoichiometry) molarity (0.0100 M) by volume of titrant (0.02728 L).

Mol Sn^{2+} = (0.02728 L MnO_4^-)$\left(\frac{0.0100\ \text{mol MnO}_4^-}{1\ \text{L}}\right)\left(\frac{5\ \text{mol Sn}^{2+}}{2\ \text{mol MnO}_4^-}\right)$ = 6.82×10^{-4} mol Sn^{2+}

This number of moles of tin ion remaining is subtracted from the total that was available in the 50.0 mL portion that was titrated,

Mol Sn^{2+} remaining = 0.0015 mol Sn^{2+} – 6.82×10^{-4} mol Sn^{2+} = 8.18×10^{-4} mol Sn^{2+} and the answer is converted to the number of moles of MnO_4^- that had reacted with this number of moles of Sn^{2+}.

Mol MnO_4^- = (8.18×10^{-4} mol Sn^{2+})$\left(\frac{2\,\text{mol MnO}_4^-}{5\,\text{mol Sn}^{2+}}\right)$ = 3.27×10^{-4} mol MnO_4^-

Multiply this number by 10 to get the moles of MnO_4^- that had not reacted with the SO_2 in the original 500 mL of 0.0200 M $KMnO_4$.

3.27×10^{-4} mol $MnO_4^- \times 10 = 3.27 \times 10^{-3}$ mol MnO_4^-

By difference, calculate the moles of SO_2 that had reacted, which is equal to the number of moles of S in the original sample.

Mol MnO_4^- added to SO_2 = (500 mL MnO_4^-)$\left(\frac{0.0200\ \text{mol MnO}_4^-}{1000\ \text{mL}}\right)$ = 0.0100 mol MnO_4^-

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Balanced reaction of SO_2 and MnO_4^-

$SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$

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

$5SO_2 + 2\,MnO_4^- + 2H_2O \rightarrow 2Mn^{2+} + 5SO_4^{2-} + 4H^+$

$$\text{Mol } SO_2 = (0.0100 \text{ mol } MnO_4^- - 3.27 \times 10^{-3} \text{ mol } MnO_4^-)\left(\frac{5 \text{ mol } SO_2}{2 \text{ mol } MnO_4^-}\right) = 0.0168 \text{ mol } SO_2$$

The mass of S is calculated by dividing moles by atomic mass,

$$\text{g } SO_2 = (0.0168 \text{ mol } SO_2)\left(\frac{1 \text{ mol S}}{1 \text{ mol } SO_2}\right)\left(\frac{32.067 \text{ g S}}{1 \text{ mol S}}\right) = 0.540 \text{ g S}$$

And the percentage of S in the original sample is the mass of S divided by the total sample mass, times 100.

$$\text{Percentage of S} = \frac{0.540\,\text{g S}}{1.045\,\text{g sample}} \times 100\% = 51.7\%\ \text{S}$$

5. How many mL of 0.05 M $KMnO_4$ is required to oxidize 2.0 g of $FeSO_4$ in dilute solution?

Solution

The redox reactions are

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

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

Therefore, the overall reaction is

$$MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$$

Now, by the relation $M_1V_1 = M_2V_2$, we have

$$5 \times 0.05 \times V = \frac{2}{152} \times 1000$$

Solving, we get $V = 52.63$ mL.

6. Metallic tin in the presence of HCl is oxidized by $K_2Cr_2O_7$ to stannic chloride. What volume of decinormal dichromate solution will be reduced by 1 g of tin? (Atomic weight of Sn = 118.7 u)

Solution

Balanced chemical equation is

$$2K_2Cr_2O_7 + 3Sn + 28HCl \rightarrow 4KCl + 4CrCl_3 + 3SnCl_4 + 14H_2O$$

$$2(2\times39 + 2\times52 + 7\times16) \quad 3\times118.7$$

$$= 2\times294 = 588\,\text{g} \quad = 356.1\,\text{g}$$

Now 356.1 g Sn react with $K_2Cr_2O_7$ = 588 g

1 g Sn will react with $K_2Cr_2O_7$ = $\frac{588}{356.1}\,\text{g} = 1.651\,\text{g}$

Equivalent weight of $K_2Cr_2O_7$ = $\frac{\text{Molecular weight}}{6} = \frac{294}{6} = 49$

Decinormal $K_2Cr_2O_7$ solution means 1000 mL solution contains 4.9 g $K_2Cr_2O_7$, that is, 4.9 g $K_2Cr_2O_7$ are present in 1000 mL solution.

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Thus, 1.651 g $K_2Cr_2O_7$ will be present in $= \frac{1000}{4.9} \times 1.651 = 336.9\,\text{mL}$

7. If 2.68×10^{-3} mol of a solution containing an ion A^{n+} requires 1.61×10^{-3} mol of MnO_4^- for the oxidation of A^{n+} to AO_3^- in acid medium, what is the value of *n*?

Solution

Since in acidic medium, A^{n+}, is oxidized to AO_3^-, the change in oxidation state from (+5) to (+*n*) = 5 – *n*

Total number of electrons that have been given out during oxidation of 2.68×10^{-3} mol of $A^{n+} = 2.68 \times 10^{-3} \times (5 - n)$

Thus, the number of electrons added to reduce 1.61×10^{-3} mol of MnO_4^- to Mn^{2+}, that is, (+7) to (+2) = $1.61 \times 10^{-3} \times 5$

Therefore,

$$1.61 \times 10^{-3} \times 5 = 2.68 \times 10^{-3} \times (5 - n)$$

Solving, we get $n = 2$.

8. Both $Cr_2O_7^{2-}(aq)$ and $MnO_4^-(aq)$ can be used to titrate $Fe^{2+}(aq)$. If in a given titration, 24.50 cm^3 of 0.1 M $Cr_2O_7^{2-}$ were used, then what volume of 0.1 M MnO_4^- solution would have been used for the same titration?

Solution

The balanced redox reactions involving oxidation of Fe^{2+} ions by $Cr_2O_7^{2-}$ and MnO_4^-, respectively, are

$$Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$$

$$MnO_4^- + +5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$$

Suppose V_2 cm^3 of M_2 Fe^{2+} is titrated against 24.50 cm^3 of 0.1 M $Cr_2O_7^{2-}$ and V_1 cm^3 of 0.1 M MnO_4^- solutions, then

$$\frac{24.5 \times 0.1}{1}(Cr_2O_7^{2-}) = \frac{M_2V_2}{6}(Fe^{2+}) \quad (1)$$

$$\frac{V_1 \times 0.1}{1}(MnO_4^-) = \frac{M_2V_2}{5}(Fe^{2+}) \quad (2)$$

Equating Eqs. (1) and (2), we get

$$V_1 = 24.5 \times \frac{6}{5} = 29.4\ cm^3$$

9. Ceric ammonium sulphate and potassium permanganate are used as oxidizing agents in the acidic medium for oxidation of ferrous ammonium sulphate to ferric sulphate. What is the ratio of the number of moles of cerium ammonium sulphate required per mole of ferrous ammonium sulphate to the number of moles of $KMnO_4$ required per mole of ferrous ammonium sulphate?

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

Consider $NH_4)_2Ce(SO_4)_2 = Ce^{4+}$ and $(NH_4)_2FeSO_4 \cdot 6H_2O = Fe^{2+}$

Now,

$$Ce^{4+} + e^- \rightarrow Ce^{3+}$$

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

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Since reduction of $KMnO_4$ is a $5e^-$ change and oxidation of $(NH_4)_2SO_4FeSO_4{\cdot}6H_2O$ is a $1e^-$ change, oxidation of 1 mol of $(NH_4)_2SO_4FeSO_4{\cdot}6H_2O$ will require 1/5 mol of $KMnO_4$. Further, since reduction of $(NH_4)_2SO_4Ce(SO_4)_2$ is a $1e^-$ change and that of oxidation of $(NH_4)_2SO_4FeSO_4{\cdot}6H_2O$ is also a $1e^-$ change, oxidation of 1 mol of $(NH_4)_2SO_4FeSO_4{\cdot}6H_2O$ will require 1 mol of $(NH_4)_2SO_4Ce(SO_4)_2$. Thus, the ratio of the number of moles of $(NH_4)_2SO_4Ce_2(SO_4)_2$ and $KMnO_4$ required to oxidize 1 mol of $(NH_4)_2SO_4FeSO_4{\cdot}6H_2O$ is 1:1/5 or 5:1.

10. What is the number of moles of MnO_4^- required to oxidize 1 mol of ferrous oxalate completely in acidic medium?

Solution

The reaction is as follows:

$$MnO_4^- + FeC_2O_4 \rightarrow Mn^{2+} + Fe^{3+} + CO_2$$

$$n_1 \times \text{Number of moles of } MnO_4^- = n_2 \times \text{Number of moles of } FeC_2O_4$$

Therefore, $n_1 \times 5 = 3 \times 1 = 0.6 \text{ mol}$.

11. An impure sample of sodium oxalate ($Na_2C_2O_4$) weighing 0.20 g is dissolved in aqueous solution of H_2SO_4 and solution is titrated at 70°C, requiring 45 mL of 0.02 M $KMnO_4$ solution. The end point is over-run, and back titration is carried out with 10 mL of 0.1 M oxalic acid solution. Find the % purity of $Na_2C_2O_4$ in sample.

Solution

$$C_2O_4^{2-} \rightarrow 2CO_2 + 2e^- \times 5$$

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

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

10 mL of 0.1 M oxalic acid solution

$$\frac{0.1 \text{ mol}}{1000 \text{ mL}} \times 10 \text{ mL} = 10^{-3} \text{ mol of } C_2O_4^-$$

$$\frac{2 \text{ mol } MnO_4^-}{5 \text{ mol } C_2O_4^-} \times 10^{-3} \text{ mol } C_2O_4^{2-} = 4 \times 10^{-4} \text{ mol of } MnO_4^-$$

45 mL of 0.02 M $KMnO_4$

$$\frac{0.02}{1000} \times 45 = 9 \times 10^{-4} \text{ mol of } MnO_4^-$$

$$\text{Number of moles of } MnO_4^- \text{ consumed by } Na_2C_2O_4 = 9 \times 10^{-4} - 4 \times 10^{-4} = 5 \times 10^{-4} \text{ mol}$$

$$\frac{5}{2} \times 5 \times 10^{-4} = 12.5 \times 10^{-4} \text{ mol of } C_2O_4^{2-}$$

$$= 12.5 \times 10^{-4} \text{ mol of } Na_2C_2O_4$$

$$= 12.5 \times 10^{-4} \times (46 + 28 + 64) \text{ g}$$

$$= 12.5 \times 10^{-4} \times 134 \text{ g} = 0.1675 \text{ g}$$

Therefore, $$\%\text{purity} = \frac{0.1675 \text{ g}}{0.20 \text{ g}} \times 100 = 83.75\%$$

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12. A mixture of FeO and Fe_2O_3 is completely reacted with 100 mL of 0.25 M acidified $KMnO_4$ solution. The resultant solution was then titrated with Zn dust, which converted Fe^{3+} of the solution to Fe^{2+}. The Fe^{2+} required 1000 mL of 0.10 M $K_2Cr_2O_7$ solution. Find out the mass % Fe_2O_3 in the mixture.

Solution

(A) FeO and Fe_2O_3 will react with 100 mL of 0.25 M $KMnO_4$.

$$\frac{0.25}{1000}\times 100 = 25\times 10^{-3} \text{ mol of } MnO_4^-$$

$$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$$

$$\frac{5}{1}\times 25\times 10^{-3} \text{ mol} = 125\times 10^{-3} \text{ mol of } Fe^{2+}$$

$$= 125\times 10^{-3} \text{ mol of } FeO = 9 \text{ g of } FeO$$

$$= 125\times 10^{-3}\times (55.8+16) = 9 \text{ g of } FeO$$

0.1 M $K_2Cr_2O_7$ of 1000 mL

$$\frac{0.1 \text{ mol}}{1000 \text{ mL}}\times 1000 \text{ mL} = 0.1 \text{ mol of } Cr_2O_7^{2-}$$

$$Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$$

$$\frac{6}{1}\times 0.1 = 0.6 \text{ mol of } Fe^{2+}$$

$$\text{Number of moles of } Fe^{3+} = (0.6 - 0.125) \text{ mol}$$

$$= 0.475 \text{ mol of } Fe^{3+}$$

$$= 0.2375 \text{ mol of } Fe_2O_3 = 38 \text{ g of } Fe_2O_3$$

As $FeO + Fe_2O_3 = 9 + 38 = 47$ g, so the mass% is

$$\% \text{ of } Fe_2O_3 = \frac{38 \text{ g}}{47 \text{ g}}\times 100 = 80.85\%$$

13. A 150 mL of solution of I_2 is divided into two unequal parts. First part reacts with hypo solution in acidic medium. 15 mL of 0.4 M hypo was consumed. Second part was added with 100 mL of 0.3 M NaOH solution. Residual base required 10 mL of 0.3 M $H_2C_2O_4$ solution for complete neutralization. What was the initial concentration of I_2?

Solution

The reactions involved are

$$2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$$

$$I_2 + 2e^- \rightarrow 2I^-$$

$$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$$

$$I_2 + 2e^- \rightarrow 2I^- \times 5$$

$$I_2 + 6H_2O \rightarrow 2IO_3^- + 12H^+ + 10e^-$$

$$6I_2 + 6H_2O \rightarrow 10I^- + 2IO_3^- + 12H^+$$

Adding $12\bar{O}H$ on both sides of above equation, we get

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$$6I_2 + 6H_2O + 12\bar{O}H \rightarrow 10I^- + 2IO_3^- + 12H^+ + 12\bar{O}H$$

$$6I_2 + 12\bar{O}H \rightarrow 10I^- + 2IO_3^- + 6H_2O$$

$$3I_2 + 6\bar{O}H \rightarrow 5I^- + IO_3^- + 3H_2O$$

$$2NaOH + H_2C_2O_4 \rightarrow Na_2C_2O_4 + 2H_2O$$

10 mL 0.3 M $H_2C_2O_4$ means

$$\frac{0.3}{1000} \times 10 = 3 \times 10^{-3} \text{ mol of } H_2C_2O_4 = 0.3 \times 10^{-2} \text{ mol of } H_2C_2O_4$$

$$\text{Number of moles of NaOH (excess) neutralized by } H_2C_2O_4 = 0.6 \times 10^{-2} \text{ mol}$$

$$\frac{0.3 \text{ mol NaOH}}{1000 \text{ mL}} \times 100 \text{ mL} = 3 \times 10^{-2} \text{ mol of NaOH (total)}$$

$$\text{Therefore, the number of moles of NaOH reacted with } I_2 = 3 \times 10^{-2} - 0.6 \times 10^{-2} = 2.4 \times 10^{-2} \text{ mol}$$

$$\frac{3I_2}{6\bar{O}H} \times 2.4 \times 10^{-2} \text{ mol } \bar{O}H = 1.2 \times 10^{-2} \text{ mol of } I_2$$

15 mL 0.4 M hypo solution.

$$\frac{0.4 \text{ mol}}{1000 \text{ mL}} \times 15 \text{ mL} = 6 \times 10^{-3} \text{ mol of } S_2O_3^{2-}$$

$$\text{Number of moles of } I_2 = \frac{6 \times 10^{-3} \text{ mol}}{2} = 3 \times 10^{-3} \text{ mol of } I_2$$

$$\text{Total number of moles of } I_2 = 12 \times 10^{-3} + 3 \times 10^{-3} = 15 \times 10^{-3} \text{ mol}$$

$$\text{So, Concentration of } I_2 = \frac{\text{Number of moles}}{\text{Liters of solution}} = \frac{15 \times 10^{-3}}{0.150} = \frac{15 \times 10^{-3}}{15 \times 10^{-2}} = 0.1 \text{ M}$$

14. One gram of a sample of $CaCO_3$ was strongly heated and the CO_2 liberated was absorbed in 100 mL of 0.5 M NaOH. Assuming 90% purity for the sample, how much mL of 0.5 M HCl would be required to react with the solution of the alkali and $NaCO_3$ to reach the phenolphthalein end point?

Solution

For 1 g sample of $CaCO_3$, %purity = 90%, so

$$1 \text{ g} = \frac{90}{100} = 0.9 \text{ g of } CaCO_3$$

For the reactions, $\underset{100 \text{ g}}{CaCO_3} \xrightarrow{\Delta} \underset{56 \text{ g}}{CaO} + \underset{44 \text{ g}}{CO_2}$

$$\frac{44 \text{ g}}{100 \text{ g}} \times 0.9 \text{ g} = 0.396 \text{ g of } CO_2$$

$$\text{Moles of } CO_2 = \frac{0.396 \text{ g}}{44 \text{ g}} = 9 \times 10^{-3} \text{ mol of } CO_2$$

Given that 100 mL of 0.5 M NaOH are used, so

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$$\frac{0.5\text{ mol}}{1000\text{ mL}} \times 100\text{ mL} = 50 \times 10^{-3}\text{ mol of NaOH}$$

$$\underset{9\times10^{-3}\text{ mol}}{CO_2} + 2NaOH \rightarrow Na_2CO_3 + H_2O$$

Therefore, 18×10^{-3} mol of NaOH is consumed by CO_2.

$$\text{Remaining NaOH} = 50\times10^{-3} - 18\times10^{-3} = 32\times10^{-3}\text{ mol}$$

$$\text{Number of moles of } Na_2CO_3 \text{ formed} = 9\times10^{-3}\text{ mol}$$

$$NaOH + HCl \rightarrow NaCl + H_2O$$

$$Na_2CO_3 + 2HCl \rightarrow H_2CO_3 + 2NaCl$$

Number of moles of HCl for NaOH	=	32×10^{-3} mol
Similarly, number of moles of HCl for Na_2CO_3	=	18×10^{-3} mol
Total number of moles of HCl	=	50×10^{-3} mol

Given that 0.5 M HCl is used, so

$$\text{mL of HCl} = \frac{1000}{0.5} \times 50\times10^{-3} = 100\text{ mL}$$

15. The concentration of bivalent lead ions in a sample of polluted water that also contains nitrate ions is determined by adding solid sodium sulphate (M_B = 142 g mol^{-1}) to exactly 500 mL of water. Calculate the molarity of lead ions if 0.355 g of sodium sulphate was needed for complete precipitation of lead ions as sulphate.

Solution

$$\underset{\substack{0.355\text{ g}\\2.5\times10^{-3}\text{ mol}}}{Na_2SO_4} + Pb(NO_3)_2 \rightarrow PbSO_4 + 2NaNO_3$$

$$\text{So, the number of moles of } Na_2SO_4 = \frac{0.355\text{ g}}{142} = 2.5\times10^{-3}\text{ mol of } Na_2SO_4$$

Thus, molarity of lead ions is

$$\frac{2.5\times10^{-3}}{0.5} = 0.005\text{ M} = 5\times10^{-3}\text{ M}$$

16. One liter of a mixture of O_2 and O_3 at NTP were allowed to react with an excess of acidified solution of KI. The iodine liberated required 40 mL of M/10 sodium thiosulphate solution for titration. Knowing that ultraviolet radiations of wavelength 300 nm can decompose ozone and assuming that one photon can decompose one molecule of ozone, what is the number of photons required for the complete decomposition of ozone in the original mixture?

Solution

The reactions involved are

$$O_3 \rightarrow O_2 + O$$

$$2KI + H_2SO_4 + [O] \rightarrow K_2SO_4 + H_2O + I_2$$

$$2Na_2SO_3 + I_2 \rightarrow Na_2S_4O_6 + 2NaI$$

$$2Na_2SO_3 + 2KI + O_3 + H_2SO_4 \rightarrow O_2 + K_2SO_4 + H_2O + Na_2S_4O_6 + 2NaI$$

2 mol of Na_2SO_3 react with 1 mol of O_3.

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Number of moles of Na_2SO_3 in 40 mL of M/10 Na_2SO_3 = $\frac{1}{10}\times\frac{1}{1000}\times 40 = 4\times10^{-3}$ mol

Therefore, O_3 reacted (present) = 2×10^{-3} mol

Number of molecules in 2×10^{-3} mol of O_3 = $2 \times 10^{-3} \times 6.022 \times 10^{23} = 1.2044 \times 10^{21}$

Therefore, the number of photons required = 1.2044×10^{21}

17. Gastric juice contains about 3.0 g of HCl per liter. If a person produces about 2.5 L of gastric juice per day, how many antacid tablets each containing 400 mg of $Al(OH)_3$ are needed to neutralize all the HCl produced in one day?

Solution

HCl is neutralized by $Al(OH)_3$ according to the equation

$$Al(OH)_3 + 3HCl \longrightarrow AlCl_3 + 3H_2O$$

Mass of $Al(OH)_3$ = 27 + 3 × 17 = 78 g, and mass of HCl = 3 × 36.5 = 109.5 g

HCl produced from 2.5 liter of gastric juice = 3.0 × 2.5 = 7.5 g

109.5 g HCl is neutralized by $Al(OH)_3$ = 78 g

$$7.5 \text{ g HCl will be neutralized by } Al(OH)_3 = \frac{78 \text{ g}}{109.5 \text{ g}} \times 7.5 \text{ g} = 5.342 \text{ g} = 5342 \text{ mg}$$

$$\text{Number of tablets required} = \frac{5342}{400} = 13.35 \approx 14 \text{ tablets}$$

18. 20 mL of a solution containing 0.2 g of impure sample of H_2O_2 reacts with 0.316 g of $KMnO_4$ in presence of H_2SO_4 as per following reaction.

$$KMnO_4 + H_2O_2 + H_2SO_4 \rightarrow MnSO_4 + O_2 + H_2O$$

(A) Find the purity of H_2O_2 solution.

(B) Calculate the volume of dry oxygen evolved at 27°C and at 750 mm Hg.

Solution

First write the complete balanced reaction:

$$2KMnO_4 + 5H_2O_2 + 3H_2SO_4 \rightarrow 2MnSO_4 + K_2SO_4 + 5O_2 + 8H_2O$$

(A) From the balanced reaction, we have

5 mol of $H_2O_2 \equiv$ 2 mol of $KMnO_4$

Moles of $KMnO_4$ = $\frac{0.316}{158} = 0.002$ mol

So, Moles of H_2O_2 required = $\frac{5}{2}\times\frac{0.316}{158}$

Grams of H_2O_2 = $\frac{5}{2}\times\frac{0.316}{158}\times 34 = 0.17$ g

Thus, % purity of H_2O_2 = $\frac{0.17}{0.2}\times 100 = 85\%$

(B) Now, 2 mol of $KMnO_4 \equiv$ 5 mol of O_2

$$\text{Moles of } O_2 \text{ produced} = \frac{5}{2}\times\frac{0.316}{158} = 5\times10^{-3} \text{ mol}$$

Using $pV = nRT$, we have

$$V = \frac{nRT}{p} = \frac{5\times10^{-3}\times0.0821\times300}{750/760} = 0.124 \text{ L} = 124 \text{ mL}$$

19. Borax in water gives:

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$B_4O_7^{2-} + 7H_2O \rightarrow 4H_3BO_3 + 2OH^-$

How many grams of borax ($Na_2B_4O_7 \cdot 10H_2O$) are required to:

(A) Prepare 50 mL of 0.2 M solution.

(B) Neutralize 25 mL of 0.1934 M of HCl and H_2SO_4 separately.

Solution

(A) We know that

$$\text{Molarity} = \frac{\frac{w}{M_B}}{\text{Volume (mL)}} \times 1000$$

Substituting the given data, we have

$$0.2 = \frac{\frac{w}{381.87}}{50} \times 1000 \Rightarrow w = \frac{0.2 \times 381.87}{20} = 3.82 \text{ g}$$

(B) For neutralization of HCl, we have

Milliequivalent of HCl = Milliequivalent of borax

$$25 \times 0.1934 = \frac{w}{382/2} \times 1000$$

Therefore, weight of borax = 0.09235 g.

For neutralization of H_2SO_4,

Milliequivalent of borax = Milliequivalent of H_2SO_4

$$\frac{w}{382/2} \times 1000 = 25 \times 0.1934 \times 1$$

Therefore, weight of borax = 1.847 g.

20. Consider 2.36 g of sample of dolomite containing only $CaCO_3$ and $MgCO_3$ to be dissolved in 700 mL of 0.1 N HCl. The solution was diluted to 2.5 L; 25 mL of this solution required 20 mL of 0.01 N NaOH solution for complete neutralization. Find the percentage composition of the ore.

Solution

When 2.36 g ($CaCO_3$ and $MgCO_3$) is added to 700 mL of 0.1 N HCl and diluted to a 2.5 L solution, we have

25 mL of above (diluted) sample ≡ 20 mL of 0.01 N NaOH

Note that sample contains $CaCO_3$ and $MgCO_3$, which are bases, so both will be neutralized. (Both of these carbonates react as diacidic bases.) The diluted sample contains excess of HCl solution.

Milliequivalent of excess HCl in 25 mL = Milliequivalent of NaOH $= 0.01 \times 20 = 0.2$

or Milliequivalent of excess HCl in 2.5 L $= 0.2 \times \frac{2500}{25} = 20$

Now, Milliequivalent of HCl (added to dolomite) = $0.1 \times 700 = 70$

or Milliequivalent of HCl used for dolomite = 70 – 20 = 50

Let x g of $CaCO_3$ and $(2.36 - x)$ g of $MgCO_3$ be there in dolomite. Thus,

Mass equivalent of $CaCO_3$ + Mass equivalent of $MgCO_3 = 50$

or $$\left(\frac{x}{100/2} \times 1000\right) + \left(\frac{2.36 - x}{84/2} \times 1000\right) = 50$$

Solving, we get $20x + 56.2 - 23.8x = 50 \Rightarrow 3.8x = 6.2 \Rightarrow x = 1.63 \text{ g} = \text{Mass of } CaCO_3$

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Thus, $(2.36 - x) = 0.73$ g = mass of $MgCO_3$.
Now, percentage composition of $CaCO_3$ and $MgCO_3$ are:

$$\% \ CaCO_3 = 69\% \text{ and } \% \ MgCO_3 = 31\%$$

21. When 1.249 g of a sample of pure $BaCO_3$ and impure $CaCO_3$ containing some CaO was treated with dilute HCl, it evolved 168 mL of CO_2 at NTP. From this solution, $BaCrO_4$ was precipitated, filtered, and washed. The precipitate was dissolved in dilute sulphuric acid and diluted to 100 mL. When 10 mL of this solution was treated with KI solution, iodine was liberated, which required exactly 20 mL of 0.05 N $Na_2S_2O_3$. Calculate the percentage of CaO in the sample.

Solution

From the given data, we have

$$n_{CaCO_3} + n_{BaCO_3} = \frac{168}{22400} = 7.5 \times 10^{-3} \quad (1)$$

The reaction sequence involved is

$$2BaCO_3 \rightarrow 2BaCrO_4 \rightarrow BaCr_2O_7 \rightarrow I_2 + Na_2S_2O_3$$

Therefore,
Gram equivalent of $Na_2S_2O_3$ = Gram equivalent of I_2 = Gram equivalent of $BaCr_2O_7$

$$= \frac{20 \times 10^{-3} \times 0.05 \times 100}{10} = 1 \times 10^{-2}$$

Now, $$\text{Number of moles of } BaCr_2O_7 = \frac{1}{6} \times 10^{-2} \text{ mol}$$

and $$\text{Number of moles of } BaCrO_4 = \frac{1}{3} \times 10^{-2} = 3.33 \times 10^{-3} \text{ mol} \quad (2)$$

So, number of moles of $BaCO_3 = 3.33 \times 10^{-3}$ mol
Weight of $BaCO_3 = 0.0033 \times 197 = 0.65$ g. Hence, from Eqs. (1) and (2), we get
Moles of $CaCO_3 = (7.5 - 3.33) \times 10^{-3} = 4.17 \times 10^{-3}$ mol
Weight of $CaCO_3 = 100 \times 4.17 \times 10^{-3} = 0.417$ g
Weight of CaO = 1.249 – 0.656 – 0.417 = 0.176

$$\% \text{ of CaO} = \frac{0.176}{1.249} \times 100 = 14.09\%$$

Additional Objective Questions

Single Correct Choice Type

1. In the reaction, $Na_2S_2O_3 + I_2 \rightarrow Na_2S_4O_6 + NaI$ (not balanced), which of the following is/are true?
(I) $Na_2S_2O_3$ is a reducing agent.
(II) Iodine is an oxidizing agent.
(III) It is an example of intermolecular redox reaction.
(IV) In $Na_2S_4O_6$, the average oxidation state of S is (+5/2)
(A) I, II, IV (B) I, II (C) I, II, III (D) I, II, III, IV

Solution

(D) $\overset{+2}{S} \rightarrow \overset{+2.5}{S}; I^0 \rightarrow I^{-1}$

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2. The number of electrons involved in the reduction of nitrate ion (NO_3^-) to hydrazine (N_2H_4) is (*n*-factor)

(A) 8 (B) 7 (C) 5 (D) 3

Solution

(B) $2NO_3^- + 7e^- + 10H^+ \rightarrow N_2H_4 + 3H_2O$

3. SO_2 under atmospheric condition changes to SO_x^{2-}. If oxidation number of S in is +6, what is the value of *x* in SO_x^{2-}?

(A) 1 (B) 2 (C) 3 (D) 4

Solution

(D) $SO_4^{2-} \Rightarrow +6 + x(-2) = -2 \Rightarrow x = 4$.

4. $NaIO_3$ reacts with $NaHSO_3$ according to equation: $IO_3^- + 3HSO_3^- \rightarrow I^- + 3H^+ + 3SO_4^{2-}$. The weight of $NaHSO_3$ required to react with 100 mL of solution containing 0.66 g of $NaIO_3$ is (*n*-factor)

(A) 5.2 g (B) 4.57 g (C) 2.3 g (D) 10.4 g

Solution

(D)

In the reaction

$I^{5+} + 6e \rightarrow I^-$

$S^{6+} + 2e \rightarrow S^{4+}$

Therefore,

$$M_{equiv}\,NaHSO_3 = M_{equiv}\,NaIO_3$$

$$N_1V_1(NaHSO_3) = N_2V_2(NaIO_3)$$

$$\frac{w}{104/2} \times 100 = \frac{0.66}{198/6} \times 100$$

$$w = 1.04\text{g}$$

5. In the reaction $4I^- + Hg^{2+} \rightarrow HgI_4^{2-}$, 1 mol each of Hg^{2+} and I^- will form:

(A) 1 mol (B) 0.5 mol (C) 0.25 mol (D) 2 mol

Solution

(C) 1 mol each of Hg^{2+} and I^- will form 1/4 = 0.25 mol

6. A 0.46 g sample of As_2O_3 required 25.0 mL of $KMnO_4$ solution for its titration. The molarity of $KMnO_4$ solution is

(A) 0.016 (B) 0.074 (C) 0.032 (D) 0.128

Solution

(B) The reaction involved is $5As_2O_3 + 4MnO_4^- + 9H_2O + 12H^+ \rightarrow 10HAsO_4^{3-} + 4Mn^{2+}$

5 mol of As_2O_3 reacts with 4 mol of MnO_4^-

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Therefore, $0.46/198$ mol of As_2O_3 will reacts with $\frac{4}{5}\times\frac{0.46}{198}=0.00186$ mol of MnO_4^-

Hence, molarity of $KMnO_4$ = $\frac{0.00186}{25}\times 1000 = 0.074$ M

7. Number of moles of $K_2Cr_2O_7$ reduced by 1 mol of Sn^{2+} is
(A) 1/3 (B) 1/6 (C) 2/3 (D) 1
Solution
(A) $Cr_2O_7^{2-}+3Sn^{2+}+14H^+\rightarrow 3Sn^{4+}+2Cr^{3+}+7H_2O$
3 mol of Sn^{2+} react with 1 mol of $Cr_2O_7^{2-}$.
So, 1 mol of Sn^{2+} will react with (1/3) mole of $Cr_2O_7^{2-}$.

8. Equivalent weight of H_3PO_2 (molecular weight = *M*) when it disproportionates into PH_3 and H_3PO_3 is
(A) *M* (B) *M*/2 (C) *M*/4 (D) 3*M*/4
Solution
(D) The reaction involved is

$$3H_3\overset{+1}{P}O_2 \rightleftharpoons \overset{-3}{P}H_3 + 2H_3\overset{+3}{P}O_3$$

The number of electrons transferred in oxidation half reaction is 4 and in the reduction half reaction is 2. Hence, the effective number of electrons transferred = 4, and total number of moles involved = 3.

Therefore, Equivalent weight = $\frac{\text{Molecular weight}}{n\text{-factor}}=\frac{M}{4/3}=\frac{3M}{4}$

9. Amongst the following, identify the species with an atom in +6 oxidation state.
(A) MnO_4^- (B) $Cr(CN)_6^{3-}$ (C) NiF_6^{2-} (D) CrO_2Cl_2
Solution
(D) Let *x* be the oxidation state of Cr. Then $x + 2(-2) + 2(-1) = 0 \Rightarrow x = +6$. Similary for the others the oxidation states are $\overset{+7}{Mn}O_4^-$, $\overset{+3}{Cr}(CN)_6^{3-}$, $\overset{+4}{Ni}F_6^{2-}$.

10. The reaction $3ClO^-(aq) \rightarrow ClO_3^-(aq) + 2Cl^-(aq)$ is an example of
(A) Oxidation reaction (B) Reduction reaction
(C) disproportionation reaction (D) Decomposition reaction
Solution
(C) $3\overset{+1}{Cl}O^-(aq) \rightarrow \overset{+5}{Cl}O_3^-(aq) + 2\overset{-1}{Cl}^-(aq)$

11. Which oxyacid of Cl cannot undergo disproportionation?
(A) HClO (B) $HClO_2$ (C) $HClO_3$ (D) $HClO_4$
Solution
(C) Disproportionation is the process in which an element undergoes both oxidation and reduction in a reaction. Moreover, an element can show a maximum oxidation state equal to its "group number" and a minimum oxidation state equal to its "group number 8". Among the oxyacids of chlorine given, in $HClO_4$ chlorine is in its highest oxidation state (+7). So, it cannot undergo further oxidation, and hence, $HClO_4$ cannot undergo disproportionation.

12. Which of the following oxides is most basic?

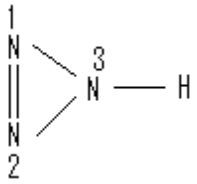
(A) MnO (B) Mn_2O_3 (C) MnO_2 (D) Mn_2O_7

Solution

(A) Less positive charge, more basic oxide.

$$\overset{+2}{M}nO, \overset{+3}{M}n_2O_3, \overset{+4}{M}nO_2, \overset{+7}{M}n_2O_7$$

13. In this compound HN_3 (hydrazoic acid), oxidation state of N^1, N^2, and N^3 are



(A) 0, 0, 3 (B) 0, 0, −1 (C) 1, 1, −3 (D) −3, −3, −3

Solution

(B) Te oxidation states are as follows:

$N^1 = 0$ $N^2 = 0$ $N^3 = -1$(as it is attached to H)

14. Which of the following agents is the most oxidizing?

(A) O_3 (B) $KMnO_4$ (C) H_2O_2 (D) $K_2Cr_2O_7$

Solution

(A) Among the given oxidizing agents, ozone is the strongest oxidizing agent because on decomposition it forms atomic oxygen.

$$O_3 \rightarrow O_2 + O$$

15. When 1 mol of $KClO_3$ takes 4 mol of electrons, the expected product is

(A) ClO^{2-} (B) ClO^{4-} (C) OCl^- (D) Cl^-

Solution

(C)

$$\overset{+5}{K}\overset{}{Cl}O_3 + 4e^- \rightarrow KO\overset{+1}{Cl}^- + O_2$$

16. On the basis of structure, the oxidation of two Cl atoms in $CaOCl_2$, respectively, are

(A) −1 and +1. (B) +2 and −2. (C) −2 and +2. (D) −1 and +3.

Solution

(A) $CaOCl_2$ has two chlorine atoms having different oxidation states.

Ca^{2+} $(OCl)^-$ Cl^- where $(OCl)^-$ is the hypochloride ion and Cl^- chloride ion. In the hypochoride ion, chlorine is in +1 state, while in chloride ion, it is in −1 state.

17. A blue color appears on addition of aqueous ammonia to aqueous $CuSO_4$ due to reaction:

$$CuSO_4 + 4NH_3 \rightarrow [Cu(NH_3)_4]SO_4$$

In this reaction,

(A) copper is oxidized. (B) copper is reduced. (C) ammonia is reduced. (D) none of these.

Solution

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(D) $\overset{2+}{Cu}\overset{2-}{S}O_4 + 4\overset{-3}{N}\overset{+1}{H_3} \rightarrow [\overset{+2}{Cu}(\overset{-3}{N}\overset{+1}{H_3})_4]SO_4$

In this reaction color appears due to *d–d* transition that takes place in copper when aqueous NH_3 (ligand) is added to the aqueous copper sulphate solution.

18. Among the following molecules, in which does bromine show the maximum oxidation number?

(A) $Hg_2(BrO_3)_2$ (B) Br^-Cl (C) $KBrO_4$ (D) Br_2

Solution

(C) The maximum oxidation state possible for Br = +7. In $KBrO_4$, Br exists in +7 oxidation state.

19. When copper is treated with a certain concentration of nitric acid, nitric oxide, and nitrogen dioxide is liberated in equal volumes according to the equation,

$$xCu + yHNO_3 \rightarrow Cu(NO_3)_2 + NO + NO_2 + H_2O$$

The coefficients of *x* and *y* are

(A) 2 and 3. (B) 2 and 6. (C) 1 and 3. (D) 3 and 8.

Solution

(B) Balanced equations for producing NO and NO_2, respectively, are

$$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O \quad (1)$$

$$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O \quad (2)$$

Adding Eqs. (1) and (2), we have

$$4Cu + 12HNO_3 \rightarrow 4Cu(NO_3)_2 + 2NO_2 + 2NO + 6H_2O$$

$$2Cu + 6HNO_3 \rightarrow 2Cu(NO_3)_2 + NO_2 + NO + 3H_2O$$

Thus, coefficients *x* and *y* of Cu and HNO_3 are 2 and 6, respectively.

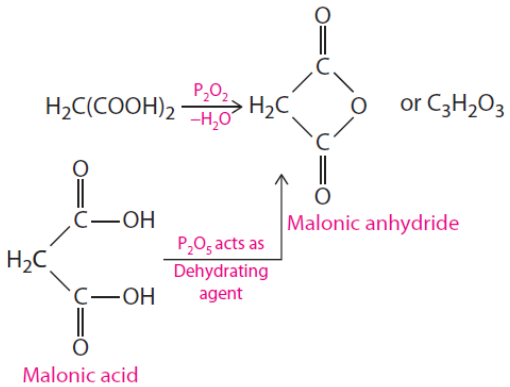
20. What is the oxidation number of carbon in A?

$$CH_2(COOH)_2 \xrightarrow{P_4O_{10}} A$$

(A) +4 (B) +3/4 (C) +4/3 (D) +2/3

Solution

(C) The reaction is



Suppose oxidation state of carbon in $C_3H_2O_3 = x$

$$3 \times x + 2 \times 1 + 3 \times (-2) = 0 \Rightarrow 3x + 2 - 6 = 0 \Rightarrow 3x - 4 = 0 \Rightarrow x = +4/3$$

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21. When iodine is passed through sodium thiosulphate, it is decolorized and sodium tetrathionate is formed. What is the change in the oxidation number of a sulphur atom in the reaction?

(A) 1/2 (B) 3/2 (C) 2 (D) 1

Solution

(A) The oxidation states are

$$\underset{\text{Sodium thiosulphate}}{Na_2\overset{+2}{S_2}O_3} \rightarrow \underset{\text{Sodium tetrathionate}}{Na_2\overset{+2.5}{S_4}O_6}$$

Change in oxidation number of sulphur = 1/2.

22. In the alkaline medium, the color of potassium dichromate solution changes from orange to yellow due to the formation of potassium chromate. What is the change in oxidation state of chromium?

(A) 1 (B) 2 (C) 3 (D) 0

Solution

(D) $\underset{\text{Potassium dichromate}}{K_2\overset{+6}{Cr_2}O_7} \xrightarrow[\text{medium}]{\text{Alkaline}} \underset{\text{Potassium chromate}}{K_2\overset{+6}{Cr}O_4}$

No change in the oxidation state of chromium takes place.

23. The oxidation number of carbon in CH_2Cl_2 is

(A) –4 (B) +4 (C) 0 (D) –2

Solution

(C) Suppose oxidation number of C = x, then $x + 2 - 2 = 0 \Rightarrow x = 0$

24. SO_2 is passed through an acidified solution of potassium dichromate, the oxidation state of S changes from

(A) +4 to 0 (B) +4 to +2 (C) +4 to +6 (D) +6 to +4

Solution

(C) The reaction can be represented as

$$\overset{+6}{Cr_2}O_7^{2-} + \overset{+4}{SO_2} \longrightarrow \overset{3+}{2Cr} + \overset{+6}{SO_4}^{2-}$$

O.N + 4 + 0 + 6

25. The oxidation states of sulphur in the anions $SO_3^{2-}, S_2O_4^{2-}$ and $S_2O_6^{2-}$ follow the order

(A) $S_2O_4^{2-} < SO_3^{2-} < S_2O_6^{2-}$ (B) $SO_3^{2-} < S_2O_4^{2-} < S_2O_6^{2-}$

(C) $S_2O_4^{2-} < S_2O_6^{2-} < SO_3^{2-}$ (D) $S_2O_6^{2-} < S_2O_4^{2-} < SO_3^{2-}$

Solution

(A) $\overset{+3}{S_2}O_4^{2-} < \overset{+4}{S}O_3^{2-} < \overset{+5}{S_2}O_6^{2-}$

26. Which of the following compounds contain the metal atom in its highest possible oxidation state?

(A) KSCN (B) MnO_2 (C) Cr_2O_3 (D) H_2SnCl_4

Solution

(A) In this compound K is present in the highest possible oxidation state.

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$\overset{+1}{K}\overset{-2}{S}\overset{+4}{C}\overset{-3}{N}$

27. A compound of Xe and F is found to have 53.5% Xe. What is the oxidation number of Xe in this compound?

(A) –4 (B) 0 (C) +4 (D) +6

Solution

(D)

Element	%	Atomic mass	Relative no. of moles	Simple ratio of moles
Xe	53.5	131.3	53.5/131.3 = 0.41	1
F	46.5	19	46.5/19 = 2.5	6

Formula of the compound = XeF_6

Hence, oxidation number of Xe = +6.

28. Which reaction involves neither oxidation nor reduction?

(A) $Cr_4^{2-} \rightarrow Cr_2O_7^{2-}$ (B) $Cr \rightarrow CrCl_3$ (C) $Na \rightarrow Na^+$ (D) $2S_2O_3^{2-} \rightarrow S_4O_6^{2-}$

Solution

(A) $\overset{+6}{C}r\overset{2-}{O_4} \rightarrow \overset{+6}{C}r_2\overset{2-}{O_7}$

No change in oxidation number.

29. Which one of the following is an example of disproportionation?

(A) $NH_3 + 3CuO \rightarrow 3Cu + 3H_2O + N_2$ (B) $5HI + HIO_3 \rightarrow 3H_2O + 3I_2$

(C) $I_2 + 2Na_2S_2O_3 \rightarrow 2NaI + Na_2S_4O_6$ (D) $P_4 + 3NaOH + 3H_2O \rightarrow 3NaH_2PO_2 + PH_3$

Solution

(D) Disproportionation is a type of reaction in which the same element is oxidized and reduced as well.

$$\overset{0}{P_4} + 2NaOH \rightarrow 3NaH_2\overset{+1}{P}O_2 + \overset{-3}{P}H_3$$

In the above reaction, P is oxidized to PH_3 and reduced to NaH_2PO_2.

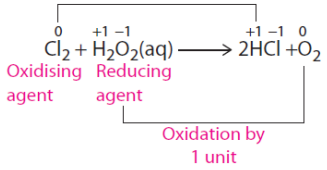
30. H_2O_2 acts as a reducing agent in

(A) $FeCl_2 + HCl + H_2O_2 \rightarrow FeCl_3 + H_2O$ (B) $Cl_2 + H_2O_2 \rightarrow 2HCl + O_2$

(C) $2HI + H_2O_2 \rightarrow I_2 + H_2O$ (D) $H_2SO_3 + H_2O_2 \rightarrow H_2SO_4 + H_2O$

Solution

(B) The reaction can be represented as



31. Which of the following has been arranged in order of increasing oxidation number of nitrogen?

(A) $NH_3 < N_2O_5 < NO < N_2$ (B) $NO_2^+ < NO_3^- < NO_2^- < N_3^-$

(C) $NH_4^+ < N_2H_4 < NH_2OH < N_2O$ (D) $NO_2 < NaN_3 < NH_4^+ < N_2O$

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Solution

(C) $\overset{-3}{N}H_4^+ \overset{-2}{N}_2H_4 < \overset{-1}{N}H_2OH < \overset{+1}{N}_2O$

32. Which of the following agents is the most reducing?

(A) Mg (B) Na (C) K (D) Br_2

Solution

(C) K is at topmost position in the electrochemical series in comparison with other elements (Mg, Na, Br_2). Therefore, it will be the most strongly reducing substance.

33. No disproportionation is possible for

(A) AsH_3 (B) SF_4 (C) H_5IO_6 (D) PCl_3

Solution

(C) The compound in which an element is present in its highest or lowest oxidation state cannot undergo disproportionation reaction. In H_5IO_6, the oxidation number of I is +7 (its highest oxidation state).

34. In the reaction

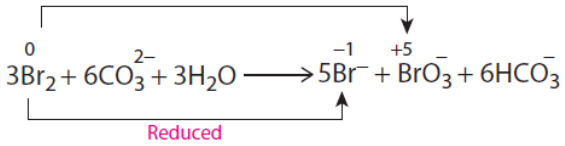
$$3Br_2 + 6CO_3^{2-} + 3H_2O \rightarrow 5Br^- + BrO_3^- + 6HCO_3^-$$

(A) bromine is oxidized and carbonate is reduced. (B) bromine is reduced and water is oxidized.

(C) bromine is neither reduced nor oxidized. (D) bromine is both reduced and oxidized.

Solution

(D) The reaction can be represented as



35. When the ion $Cr_2O_7^{2-}$ acts as an oxidant in the acidic aqueous solution, the ion Cr^{3+} is formed. How many moles of Sn^{2+} would be oxidized to Sn^{4+} by one mol of $Cr_2O_7^{2-}$ ions?

(A) 1 (B) 2 (C) 4 (D) 3

Solution

(D) The reaction can be represented as

$$CrO_7^{2-} + 6e^- + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$$

$$[Sn^{2+} \rightarrow Sn^{4+} + 2e^-] \times 3$$

$$CrO_7^{2-} + 3Sn^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 3Sn^{4+} + 7H_2O$$

36. A sample of 2.5 mol of hydrazine (N_2H_4) loses 25 mol of electrons on being converted into a new compound X. Assuming that there is no loss of nitrogen in the formation of the new compound, what is the oxidation number of nitrogen in compound X?

(A) –1 (B) –2 (C) +3 (D) +4

Solution

(C) Total oxidation of two nitrogen atoms in N_2H_4 is –4. Since it loses 25 mol electrons, the total oxidation number of two N atoms in Y increases by

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$$\underset{2.5\,\text{mol}}{\overset{-2}{N_2}H_4} \rightarrow \underset{25\,\text{mol}}{N_2^x} + \underset{10\,\text{mol}}{ne^-}$$

In hydrazine, the oxidation number of N is $2x + 4(+1) = 0$ or $x = -2$
Change in oxidation number per N atom = number of electrons lost
Hence, $2(x + 2) = 10 \Rightarrow x = +3$.

37. An element, which never has a positive oxidation state in any of its compounds is
(A) boron. (B) oxygen. (C) chlorine. (D) fluorine.
Solution
(D) This is because fluorine is the most electronegative element in the periodic table and its oxidation state is always –1.

38. The average oxidation number of sulphur in $Na_2S_4O_6$ is
(A) 1.5 (B) 2.5 (C) 3 (D) 2
Solution
(B) Suppose oxidation number of sulphur = x

$$2 \times 1 + 4 \times (x) + 6 \times (-2) = 0 \Rightarrow 2 + 4x - 12 = 0$$
$$4x - 10 = 0 \Rightarrow 4x = 10 \Rightarrow x = +2.5$$

39. Oxidation number of fluorine in F_2O is
(A) +1 (B) +2 (C) –1 (D) –2
Solution
(C) In F_2O, O is in the +2 state.

$$2x + 1(+2) = 0 \Rightarrow 2x + 2 = 0 \Rightarrow x = -1$$

40. How many moles of $KMnO_4$ are required in the acidic medium for complete oxidation of 15 mol of $FeSO_4$?
(A) 2 (B) 3 (C) 4 (D) 5
Solution
(B) The reactions involved are

$$MnO_4^- + 5e^- + 8H^+ \rightarrow Mn^{+2} + 4H_2O$$
$$[Fe^{2+} \rightarrow Fe^{3+} + e^-] \times 5$$
$$\overline{MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{+2} + 5Fe^{3+} + 4H_2O}$$

Hence,
5 mol of $FeSO_4$ are completely oxidized by $KMnO_4$ = 1 mol
15 mol of $FeSO_4$ are completely oxidized by $KMnO_4$ = $\frac{1 \times 15}{5} = 3\,\text{mol}$

41. Equivalent weight of MnO_4^- in acidic, basic and neutral media is in the ratio of
(A) 3:5:15 (B) 5:3:1 (C) 5:1:3 (D) 3:15:5.
Solution
(C) The reactions are as follows:

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$$\text{Acidic medium}: \overset{+7}{\text{M}}\text{nO}_4^- + 8\text{H}^+ + 5e^- \rightarrow \text{Mn}^{+2} + 4\text{H}_2\text{O}$$

$$\text{Basic medium}: \overset{+7}{\text{M}}\text{nO}_4^- + e^- \rightarrow \overset{+6}{\text{M}}\text{nO}_4^{2-}$$

$$\text{Neutral medium}: \overset{+7}{\text{M}}\text{nO}_4^- + 2\text{H}_2\text{O} + 3e^- \rightarrow \overset{+4}{\text{M}}\text{nO}_2 + \overset{-}{\text{O}}\text{H}$$

42. Sulphur has highest oxidation slate in

(A) H_2SO_4 (B) SO_2 (C) $Na_2S_2O_3$ (D) $Na_2S_4O_6$

Solution

(A) In H_2SO_4, sulphur has +6 oxidation state in SO_2 = +4, $Na_2S_2O_3$ = +2.

43. In which of the following compound transition metal has zero oxidation state?

(A) CrO_5 (B) $NH_2 \cdot NH_2$ (C) $NOClO_4$ (D) $Fe(CO)_5$

Solution

(D) In carbonyl compounds metal has zero oxidation state, as CO is neutral and has a zero state.

44. Consider the following reactions:

(I) $C_2O_4^{2} \rightarrow CO_2$ (II) $SO_4^{2-} \rightarrow SO_3^{2-}$ (III) $MnO_4^{2-} \rightarrow MnO_4^{+}$ (IV) $Fe^{3+} \rightarrow Fe^{2+}$

Choose the correct answer:

(A) (I) and (II) show oxidation. (B) (III) and (IV) show reduction.

(C) (I) and (III) show oxidation. (D) (III) and (IV) show oxidation.

Solution

(C) Increase in oxidation number is oxidation

(i) $\overset{+3}{\text{C}}_2\text{O}_4^{2-} \rightarrow 2\overset{+4}{\text{C}}\text{O}_2$ oxidation

(ii) $\overset{+6}{\text{S}}\text{O}_4^{2-} \rightarrow \overset{+4}{\text{S}}\text{O}_3^{2-}$ reduction

(iii) $\overset{+6}{\text{Mn}}\text{O}_4^{2-} \rightarrow \overset{+7}{\text{Mn}}\text{O}_4^{-}$ oxidation

(iv) $Fe^{3+} \rightarrow Fe^{2+}$ reduction

45. In the reaction $Cl_2 + OH^- \rightarrow Cl^- + ClO_3^- + H_2O$, chlorine is

(A) oxidized (B) reduced

(C) oxidized as well as reduced (D) neither oxidized nor reduced

Solution

(C) The reaction involved is

$$\underset{\text{Reduction}}{\overset{0}{\text{Cl}_2} + \text{OH}^- \rightarrow \overset{-1}{\text{Cl}^-} + \overset{+5}{\text{ClO}_3^-}} + \text{H}_2\text{O}$$

Oxidation

46. How many moles of electrons are needed for the reduction of each mole of Cr in the reaction,

$$CrO_5 + H_7SO_4 \rightarrow Cr_2(SO_4)_3 + H_2O + O_2$$

(A) 4 (B) 3 (C) 5 (D) 7

Solution

(B) $\overset{+6}{\text{Cr}}\text{O}_5 \rightarrow \overset{+3}{\text{Cr}}_2(\text{SO}_4)_3$

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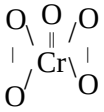

The number of moles of electron required = 3

47. Hydroxyl amine reduces iron (III) according to following equation:

$$NH_2OH + Fe_2(SO_4)_3 \rightarrow N_2(g) + H_2O + FeSO_4 + H_2SO_4$$

Which statement is correct?
(A) *n*-Factor for hydroxyl amines is 1.
(B) Equivalent weight of $Fe_2(SO_4)_3$ is M/2.
(C) 6 milliequivalent of $Fe_2(SO_4)_3$ is contained in 3 millimoles of ferric sulphate.
(D) All of these.
Solution
(D) $2\overset{-1}{N}H_2OH \rightarrow \overset{0}{N}_2$
n-factor = 1
$\overset{+3}{Fe}_2(SO_4)_3 \rightarrow 2\overset{+2}{Fe}SO_4$
n-factor = 2 (as two atoms)
Thus, equivalent weight of $Fe_2(SO_4)_3$ = Mol. wt./ 2

48. Equal volumes of 1 M each of $KMnO_4$ and $K_2Cr_2O_7$ are used to oxidize Fe(II) solution in acidic medium. The amount of Fe oxidized will be
(A) more with $KMnO_4$ (B) more with $K_2Cr_2O_7$
(C) equal with both oxidizing agents (D) cannot be determined
Solution
(B) In acidic medium,

$$MnO_1^- \rightarrow Mn^{+2} \qquad n\text{-factor} = 5/\text{mol}$$

$$Cr_2O_7^{2-} \rightarrow 2Cr^{+3} \qquad n\text{-factor} = 6/\text{mol}$$

Hence, the amount of Fe(II) oxidized is more with $Cr_2O_7^{2-}$.

49. What volume of 1 M FeC_2O_4 (ferrous oxalate) solution is required for the complete oxidation of 100 mL of 1 M $Fe(SCN)_2$ solution into Fe^{3+}, SO_4^{2-}, CO_3^{2-}, and nitrate?
(A) 0.7 L (B) 1 L (C) 0.8 L (D) 1.1 L
Solution
(D)*n*-Factor of FeC_2O_4 = 3 as $FeC_2O_4 \rightarrow Fe^{3+} + 2CO_2$
and *n*-factor of $Fe(SCN)_2$ = 33 as $Fe(SCN)_2 \rightarrow Fe^{3+} + SO_4^{2-} + CO_3^{2-} + NO_3^-$
Let the volume of FeC_2O_4 required for oxidation be x mL; then,

Milliequivalent of FeC_2O_4 = Milliequivalent of $Fe(SCN)_2$

So, $x \times 1 \times 3 = 100 \times 1 \times 33$
or $x = 1100$ mL = 1.1 L

50. A solution contains Cu^{2+} and $C_2O_4^{2-}$ ions that on titration with 1 M $KMnO_4$ consumes 10 mL of the

oxidizing agent for complete oxidation in acidic medium. The resulting solution is neutralized with Na_2CO_3, acidified with dil. CH_3COOH, and is treated with excess of KI. The liberated iodine requires 25 mL of 1 M of hypo solution, then that will be the molar ratio of Cu^{2+} to $C_2O_4^{2-}$ ions in solution?

(A) 5:2 (B) 1:2 (C) 2:1 (D) 1:1

Solution

(D) Milliequivalent of hypo = Millliequivalent of Cu^{2+}

= 25 mass equivalent of Cu^{+2} (n = 1) = 25 mol of Cu^{2+}

$$\frac{m \text{ mol of } \begin{matrix}\text{COOH}\\|\\\text{COOH}\end{matrix}}{m \text{ mol of } KMnO_4} = \frac{5}{2}$$

Therefore, m mol of oxalic acid $= \frac{5}{2} \times 10 = 25$

$$\frac{m \text{ mol of } Cu^{2+}}{m \text{ mol of oxalic acid}} = \frac{25}{25} = 1 \Rightarrow \frac{\text{Mole of } Cu^{2+}}{\text{Mole of oxalic acid}} = 1:1$$

51.The labeling on a bottle of H_2O_2 solution is 20 "vol." then the concentration of H_2O_2 in percentage strength will be:

(A) 3.03 % (B) 5% (C) 4.55% (D) 6.06 %

Solution

(D) 1 vol. = 0.303 %. So, 20 vol. = 20 × 0.303 % = 6.06%

52. A 100 mL solution of 0.1 N HCl was titrated with 0.2 N NaOH solution. The titration was discontinued after adding 30 mL of NaOH solution. The remaining titration was completed by adding 0.25 N KOH solution. The volume of KOH required for completing the titration is

(A) 70 mL (B) 32 mL (C) 35 mL (D) 16 mL

Solution

(D) Volume of HCl neutralized by NaOH = (Caustic soda) = V_1

(HCl) $N_1V_1 = N_2V_2$ (NaOH)

$0.1 \times V_1 = 0.2 \times 30 \Rightarrow V_1 = 60$ mL

So, the remaining 40 mL 0.1 N HCl is now neutralized by 0.25 N KOH

(HCl) $N_1V_1 = N_2V_2$ (KOH)

$0.1 \times 40 = 0.25 \times V_2 \Rightarrow V_2 = 16$ mL

53. 40% *w*/*V* NaCl solution (specific gravity = 1.12) is equivalent to

(A) 3.57×10^5 ppm (B) 3.57×10^6 ppm (C) 1×10^6 ppm (D) 4×10^5 ppm

Solution

(A) For 40% *w*/*V* NaCl solution, 100 mL of solution contains 40 g of NaCl.

Given density (specific gravity) = 1.12 g mL^{-1}

100 mL solution (= 112 g) contains 40 g NaCl or 112 g of solution contains 40 g of NaCl.

10^6 g of solution contains x g of NaCl.

$$x = \frac{40}{112} \times 10^6 = 3.57 \times 10^5 \text{ ppm}$$

54. Oxalic acid $(H_2C_2O_4)$ forms two series of salt $HC_2O_4^-$ and $C_2O_4^-$. If 0.9 g of oxalic acid is in 100 mL solution, $HC_2O_4^-$ and $C_2O_4^-$ have normality, respectively,

(A) 0.1 N, 00N (B) 0.1 N, 0.2 N (C) 0.2 N, 0.2 N (D) 0.2 N, 0.1 N

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Solution

(B) $H_2C_2O_4 \longrightarrow HC_2O_4^- + H^+$ (n-factor = 1)

$H_2C_2O_4 \longrightarrow C_2O_4^- + 2H^+$ (n-factor = 2)

So, $M_{H_2C_2O_4} = \dfrac{0.9/90}{100/1000} = 0.1\,\text{M}$

or $N_{HC_2O_4^-} = 1 \times 0.1\,\text{N} = 0.1\,\text{N}$ and $N_{C_2O_4^{2-}} = 2 \times 0.1\,\text{N} = 0.2\,\text{N}$

Multiple Correct Choice Type

1. $H_2C_2O_4$ and $NaHC_2O_4$ behave as acids as well as reducing agents. Which are the correct statements?

(A) Equivalent weight of $H_2C_2O_4$ and $NaHC_2O_4$ are equal to their molecular weights when acting as reducing agents

(B) Equivalent weight of $H_2C_2O_4$ and $NaHC_2O_4$ are equal to half their molecular weights when acting as reducing agents

(C) 100 mL of 1 M solution of each is neutralized by equal volumes of 1 N $Ca(OH)_2$

(D) 100 mL of 1 M solution of each is oxidized by equal volumes of 1 M $KMnO_4$

Solution

(B, C, D)

In the reduction reactions, $H_2C_2O_4$ and $NaHC_2O_4$ are converted to CO_2, so

$(C^{3+})_2 \rightarrow 2C^{4+} + 2e$

Therefore, Equivalent weight =Molecular weight /2.

On titration with 1M KMnO4,

$Mn^{7+} + 5e \rightarrow Mn^{2+}$

Therefore,

$M_{equiv} H_2C_2O_4 = M_{equiv} KMnO_4$

$100\,\text{mL} \times 1\text{M} \times 2 = V\,\text{mL} \times 1\,\text{M} \times 5$

$M_{equiv} NaHC_2O_4 = M_{equiv} KMnO_4$

$100\,\text{mL} \times 1\text{M} \times 2 = V\,\text{mL} \times 1\,\text{M} \times 5$

So 100 mL of 1M solution of each is neutralized by equal volumes of 1M $KMnO_4$.

2. For the following balanced redox reaction,

$2MnO_4^+ + 4H^+ + Br_2 = 2Mn^{+2} + 2BrO_3^+ + 2H_2O$

If the molecular weight of MnO_4^-, Br_2 be M_x, M_y, respectively, then

(A) equivalent weight of MnO_4^+ is $M_x/5$.

(B) equivalent weight of Br_2 is $M_y/10$.

(C) the n-factor ratio of Mn^{+2} to BrO_3^- is 1:1.

(D) none of these.

Solution

(A, B, C)

$\overset{+7}{2}MnO_4^- + 4H^+ + \overset{0}{Br_2} \rightarrow \overset{+2}{2}Mn^{+2} + \overset{+5}{2}BrO_3^- + 2H_2O$

n-factor for $MnO_4^- = 5$

Equivalent weight = $M_x/5$

n-factor for Br_2 = 10

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Equivalent weight = $M_y/10$
The half reactions are

$$\overset{+2}{Mn}^{2+} + 4H_2O \rightarrow \overset{+7}{MnO_4^-} + 5e^- + 8H^+] \times 2$$

$$2\overset{+5}{BrO_3^-} + 10e^- + 12H^+ \rightarrow \overset{0}{Br_2} + 6H_2O$$

Therefore, the ratio of *n*-factors is 1:1.

3. Which of the following statements about the following reactions are wrong?
$2Cu_2O(s) + Cu_2S(s) \rightarrow 6Cu(s) + SO_2g$
(A) Both Cu_2O and Cu_2S are reduced. (B) Only Cu_2S is reduced.
(C) Cu_2S is the oxidant. (D) Only Cu_2O is reduced.
Solution
(B, C, D) $2\overset{+1}{Cu_2}\overset{-2}{O}(s) + \overset{+1}{Cu_2}\overset{-2}{S}(s) \rightarrow 6\overset{0}{Cu}(s) + \overset{+4}{S}\overset{-2}{O_2}(g)$
Both Cu_2O and Cu_2S are reduced to Cu. Also, Cu_2O acts as the oxidant while Cu_2S acts as the reductant as it gets oxidized to SO_2.

4. Which of the following substances undergo disproportionation reactions under basic medium?
(A) F_2 (B) P_4 (C) S_8 (D) Br_2
Solution
(B, C, D) Except F_2 (which shows oxidation state of –1 only), all the other substances show both +ve and –ve oxidation states and hence undergo disproportionation.

5. A reducing agent in a redox reaction undergoes
(A) a decrease in oxidation number. (B) an increase in oxidation number.
(C) loss of electrons. (D) gain of electrons.
Solution
(B, C) A reducing agent in a redox reaction undergoes an increase in the oxidation number with loss of electrons.

6. Which of the following statements are correct?
(A) All reactions are oxidation and reduction reactions.
(B) Oxidizing agent is itself reduced.
(C) Oxidation and reduction always go side by side.
(D) Oxidation number during reduction decreases.
Solution
(B, C, D) Statement (A) is not correct. For example, during neutralization reactions and precipitation reactions oxidation number does not change. Hence, these reactions (collectively known as metathesis reaction) are not redox reactions.

7. The metals undergoing disproportionation are
(A) Sn (B) Na (C) Cu (D) Ca
Solution
(A, C) Sn and Cu can undergo disproportionation reaction.

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For Cu : $Cu^{+} \rightarrow Cu^{2+} + e^{-}$ For Sn : $Sn^{+2} \rightarrow Sn^{+4} + 2e^{-}$

Overall $\underline{2Cu^{+} \rightarrow Cu^{2+} + Cu}$ Overall $\underline{2Sn^{+2} \rightarrow Sn^{+4} + Sn}$

8. Which of the following is correct about acidic nature?

(A) $H_3PO_4 > H_3PO_3$ (B) $H_3AsO_4 > H_3AsO_3$ (C) $HClO_4 > HClO_3$ (D) $HNO_3 > HNO_2$

Solution

(B, C, D) In the given acid pairs, the one in which the central atom is present in its higher oxidation state is stronger acid than the other.

$\overset{+5}{H_3As}O_4 > \overset{+3}{H_3As}O_3$ $\overset{+7}{HCl}O_4 > \overset{+5}{HCl}O_3$ $\overset{+5}{HN}O_3 > \overset{+3}{HN}O_2$

9. In which of the following, oxidation number of oxygen is fractional?

(A) B_4O_{10} (B) B_2H_6 (C) CsO_2 (D) KO_3

Solution

(A, C, D) Suppose oxidation state of oxygen = x

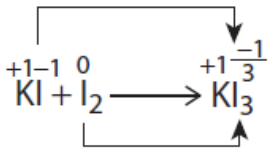
B_4O_{10}	CsO_2	KO_3
$4 \times 1 + 10\,(x) = 0$	$+1 \times 2\,(x) = 0$	$+1 \times 3\,(x) = 0$
$4 + 10x = 0$	$1 \times 2x = 0$	$1 = -3x$
$4 = -10x$	$1 = -2x$	$1 = -3x$
$x = -\frac{4}{10}$	$x = -\frac{1}{2}$	$x = -\frac{1}{3}$

10. The reaction $KI + I_2 \rightarrow KI_3$ is

(A) oxidation. (B) reduction. (C) complex formation. (D) none of these.

Solution

(A, B, C) The reaction can be represented as



11. The activity series of metals is

Au < Ag < Cu < Sn < Cd < Zn < Al < Mg < Na < Cs

Which of the following reactions does not occur spontaneously upon mixing the reagents shown?

(A) $Cd(s) + Al^{3+}(aq) \rightarrow Cd^{2+}(aq) + Al(s)$ (B) $Cd(s) + Cu^{2+}(aq) \rightarrow Cd^{2+}(aq) + Cu(s)$

(C) $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$ (D) $Al(s) + Ag^{+}(aq) \rightarrow Al^{3+}(aq) + Ag(s)$

Solution

(A, D) Because Cd is less reactive than Al and Ag is more reactive than Al.

12. Which of the following represent redox reactions?

(A) $Cr_2O_7^{2-} + 2OH^{-} \rightarrow 2CrO_4^{2-} + H_2O$

(B) $2CrO_4^{2-} + 2H^{+} \rightarrow Cr_2O_7^{2-} + H_2O$

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(C) $2MnO_4^- + 3Mn^{2+} + 4OH^- \rightarrow 5MnO_2 + 2H_2O$

(D) $IO_4^- + I^- + H^+ \rightarrow I_2 + H_2O$

Solution

(C, D) Both (C) and (D) are disproportionation reactions (redox)

Redox reactions are those in which both reduction and oxidation occur simultaneously.

13. For the following balanced redox reaction,

$$2MnO_4^- + 4H^+ + Br_2 \rightarrow 2Mn^{2+} + 2BrO_3^+ + 2H_2O$$

if the molecular weight of MnO_4^- and Br_2 be M_{B1} and M_{B2} respectively, then

(A) equivalent weight of MnO_4^- is $M_{B1}/5$.

(B) equivalent weight of Br_2 is $M_{B2}/10$.

(C) the *n*-factor ratio of Mn^{2+} to BrO_3^- is 1:1.

(D) none of these.

Solution

(A, B, C)

$$\overset{+7}{2MnO_4^-} + 4H^+ + \overset{0}{Br_2} \rightarrow \overset{+2}{2Mn^{+2}} + \overset{+5}{2BrO_3^-} + 2H_2O$$

n-factor for $MnO_4^- = 5$ (as +7 to +2)

Equivalent weight = $M_x/5$

n-factor for Br_2 = 10 (as 0 to +5). Therefore, for two atoms = 5 × 2 = 10

Equivalent weight = $M_y/10$

Ratio of *n*-factor $\overset{+2}{Mn^{2+}} \rightarrow \overset{+7}{MnO_4^-}$

$\overset{+5}{3BrO_3^-} \rightarrow \overset{0}{Br_2} = 1:1$

14. When non-stoichiometric compound $Fe_{0.95}O$ is heated in presence of oxygen, then it converts into Fe_2O_3. Which of the following statements are then correct?

(A) Equivalent weight of $Fe_{0.95}O$ is $M_B/0.5$ where M_B is molecular weight of $Fe_{0.95}O$.

(B) The number of moles of Fe^{3+} and Fe^{2+} in 1 mol $Fe_{0.95}O$ are 0.1 and 0.85, respectively.

(C) The number of moles of Fe^{3+} and Fe^{2+} in 1 mol of $Fe_{0.95}O$ are 0.85 and 0.10, respectively

(D) The % composition of Fe^{2+} and Fe^{3+} in the non-stoichiometric compound is 89.47% and 10.53. respectively.

Solution

(B, D)

$Fe_{0.95}O + O_2 \rightarrow Fe_2O_3$

Let x be the fraction of Fe^{3+} in the compound; then $Fe^{2+} = (0.95 - x)$

$$3x + (0.95 - x)2 - 2 = 0 \Rightarrow x = 0.1$$

Then moles of Fe^{2+} = 0.85 mol

So, percentage composition of Fe^{3+} = 10% and of Fe^{2+} is 90%.

15. A group of students attempted to estimate the concentration of Fe^{2+} by pipetting fixed volumes of the solution into a flask, adding an excess of dilute sulphuric acid, and then titration with a standard solution of potassium manganate (VII) from a burette. The volume of $KMnO_4$ solution required by one student was 0.2 cm^3 higher than that by the other students. Which of the following are possible explanations for this discrepancy?

(A) The titration flask was rinsed with the solution of Fe^{2+} instead to water before titration.

(B) The last drop of Fe^{2+} solution was blown from the pipette into the flask.

(C) The burette was rinsed with water instead of the solution of $KMnO_4$ before titration.

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(D) One student rinsed the burette with the solution of Fe^{2+} and others rinsed with solution of $KMnO_4$ before titration.

Solution

(A, B, C) Here, we estimate the concentration of Fe^{2+} ions; so, the options A, B, and C are the correct reasons. By rinsing the flask with solution of Fe^{2+}, the concentration of Fe^{2+} in the flask increases. If the burette is rinsed with $KMnO_4$, the concentration of KMnO4 in the burette reduces. Hence more volume of $KMnO_4$ is required for titration. Lastly, when the last drop of Fe^{2+} solution was blown from the pipette into the flask, the concentration of Fe^{2+} obtained is higher than the actual one. So, the options a, b, and c bring out discrepancies in the titration.

16. 40 g NaOH, 106 g Na_2CO_3 and 84 g $NaHCO_3$ is dissolved in water and the solution is made 1 L; 20 mL of this stock solution is titrated with 1 N HCl; hence, which of the following statements are correct?

(A) The titre reading of HCl will be 40 mL if phenolphthalein is used as indicator from the very beginning.

(B) The titre reading of HCl will be 60 mL if phenolphthalein is used as indicator from the very beginning.

(C) The titre reading of HCl will be 40 mL if the methyl orange is used as indicator after the first end point.

(D) The titre reading of HCl will be 80 mL if methyl orange is used as indicator from the very beginning.

Solution

(A, C, D) NaOH and $NaHCO_3$ in the solution react as

$$NaOH + NaHCO_3 \rightarrow Na_2CO_3 + H_2O$$
$$40\text{ g} \quad 84\text{ g}$$

Now, number of moles of NaOH = $\frac{40}{40} = 1\text{ mol}$ and of $NaHCO_3 = \frac{84}{84} = 1\text{ mol}$

After the reaction, the solution will have 2 mol of Na_2CO_3 in 1 L.

(i) For phenolphthalein,

$$\text{Milliequivalent of } Na_2CO_3 = \text{Milliequivalent of HCl}$$
$$20 \times 20 \times 1 = V \times 1 \Rightarrow V = 40\text{ mL}$$

(ii) For methyl orange after the first end point

$$\text{Milliequivalent of } NaHCO_3 = \text{Milliequivalent of } Na_2CO_3 = \text{Milliequivalent of HCl}$$
$$2 \times 20 \times 1 = 1 \times V \Rightarrow V = 40\text{ mL}$$

(iii) For methyl orange if used from the very beginning:

$$\text{Milliequivalent } Na_2CO_3 + \text{Milliequivalent of } NaHCO_3 = \text{Milliequivalent of HCl}$$
$$2 \times \text{Milliequivalentof } Na_2CO_3 = \text{Milliequivalent of HCl}$$
$$2 \times 20 \times 2 = 1 \times V \Rightarrow V = 80\text{ mL}$$

17. 150 mL of M/10 $Ba(MnO_4)_2$ in acidic can oxidize completely

(A) 150 mL 1 M Fe^{2+} (B) 50 mL 1 M FeC_2O_4

(C) 75 mL 1 M $C_2C_4^{2+}$ (D) 25 mL 1 M $K_2Cr_2O_7$ solution

Solution

(A, B, C, D)

$Ba(\overset{+7}{M}nO_4)_2 \rightarrow 2\overset{+2}{M}n^{2+}$ (*n*-factor = 10)

Milliequivalent of $Ba(MnO_4)_2$ in $150\text{ mL} \Rightarrow 150 \times 10 \times \frac{1}{10} = 150$ millieq.

Milliequivalent of 1M $Fe^{2+} \rightarrow Fe^{++} = 150 \times 1 = 150$m.Eq

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Milliequivalent t of 50 mL 1M $FeC_2O_4 \rightarrow Fe^{3+} + 2CO_2 = 50 \times 3 = 150$ m.Eq

Milliequivalent of 75 mL 1M $C_2O_4^{2-} \rightarrow 2CO_2 = 75 \times 2 = 150$ m.Eq

Milliequivalent of 25 mL 1M $C_2O_7^{2-} \rightarrow 2Cr^{3+} = 25 \times 6 = 150$ m.Eq

18. $\begin{matrix}\text{COOH}\\|\\\text{COOH}\end{matrix}$ and $\begin{matrix}\text{COOK}\\|\\\text{COOH}\end{matrix}$ behave as acid are well as reducing agent. Then, which of the following are the correct statements regarding $\begin{matrix}\text{COOH}\\|\\\text{COOH}\end{matrix}$ and $\begin{matrix}\text{COOK}\\|\\\text{COOH}\end{matrix}$?

(A) When behaves as reducing agent, then its equivalent weights are equal to half of its molecular weight respectively.

(B) 1000 mL of 1N solution of each is neutralized by 1000 mL 1N $Ca(OH)_2$

(C) 1000 mL of 1M solution of each is neutralized by 1000 mL of 1M $Ca(OH)_2$.

(D) 1000 mL of 1M solution of each is neutralized by 200 mL 2M of $KMnO_4$ in acidic medium.

Solution

(A, B, D) For $\begin{matrix}\text{COOH}\\|\\\text{COOH}\end{matrix}$ and $\begin{matrix}\text{COOK}\\|\\\text{COOH}\end{matrix}$

(A) When each behave as reducing agent

$(COOH)_2 \rightarrow 2CO_2$ n-factor $= 2$, Equivalent weight $= M / 2$ (as +3 to +4)

$\begin{matrix}\text{COOH}\\|\\\text{COOH}\end{matrix} \rightarrow 2CO_2$ n-factor $= 2$, Equivalent weight $= M / 2$

(B) For neutralization, milliequivalent of one reactant = Milliequivalent of other reactant can be neutralized $1000 \times 2 = 1000 \times 2$

$\begin{matrix}\text{COOK}\\|\\\text{COOH}\end{matrix}$ is monobasic.

19. 0.1 mol of MnO_4^- (in acidic medium) can:

(A) oxidize 0.5 mol of Fe^{2+} (B) oxidize 0.166 mol of $FeCr_2O_4$

(C) oxidize 0.25 mol of $C_2O_4^{2-}$ (D) oxidize 0.6 mol of $Cr_2O_7^{2-}$

Solution

(A, B, C)

In acidic medium

$\overset{+7}{Mn}O_4^- + 8H^+ + 5e^- \rightarrow \overset{+2}{Mn}^{+2} + 4H_2O$

Total number of electrons gained = 5

Total change in oxidation number of Mn = 7 – 2 = 5

Hence, equivalents of MnO_4^- = 0.1 × 5 = 0.5(A) 0.5 × 1 = 0.5

(B) 0.166 × 3 = 0.5

(C) 0.25 × 2 = 0.5

(D) 0.6 × 6 = 3.6

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20. The reaction $3ClO^-(aq) \rightarrow ClO_3^-(aq) + 2Cl^-(aq)$ is an example of

(A) Oxidation reaction (B) Reduction reaction

(C) Disproportionation reaction (D) Decomposition reaction

Solution

(A, B, C)

$$Na_2\overset{+1}{Cl}O^- \rightarrow \overset{+5}{Cl}O_3^- + 2\overset{-1}{Cl}^-$$

It is a disproportionate reaction, which induces both oxidation and reduction of the same element.

$$\underbrace{\underbrace{Cl^{+1}}_{\text{Oxidized}} + Cl^{+5}}_{\text{Reduced}} + Cl^{-1}$$

Assertion–Reasoning Type

Choose the correct option from the following:

(A) Statement 1 is True, Statement 2 is True; Statement 2 is a correct explanation for Statement 1.

(B) Statement 1 is True, Statement 2 is True; Statement 2 is NOT a correct explanation for Statement 1.

(C) Statement 1 is True; Statement 2 is False.

(D) Statement 1 is False; Statement 2 is True.

1. Statement 1: The equivalent weight of reducing agent, Fe^{2+} is 56 (atomic mass of Fe = 56).

Statement 2: Fe loses $2e^-$ to be converted into Fe^{2+}.

Solution

(D) Equivalent weight = molecular mass/number of electrons involved in reduction/oxidation

2. Statement 1: In the roasting of FeS_2 ore is converted into ferric oxide and SO_2 gas. The equivalent mass of FeS_2 is equal to molecular weight/11.

Statement 2: The *n*-factor for reducing agent is total net change in oxidation number per formula unit.

Solution

(A)

3. Statement 1: In CrO_5, there must be peroxide linkage.

Statement 2: The maximum oxidation number of an element cannot exceed its number of valence electron.

Solution

(A)

4. Statement 1: $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$ is a redox reaction.

Statement 2: Equivalent weight of $CaCO_3$ in this reaction is 50.

Solution

(D) $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$ is a non-redox reaction.

So equivalent weight of $CaCO_3 = \dfrac{\text{Molecular weight}}{2} = 100 = 50$

5. Statement 1: Fe reacts with HCl to produce H_2 gas.

Statement 2: Fe is a better reducing agent than H_2

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Solution
(A)

6. Statement 1: Hydrogen peroxide acts only as oxidizing agent.
Statement 2: All peroxides behave as oxidizing agents only.
Solution
(D) Hydrogen peroxide acts as an oxidizing as well as reducing agent. Peroxides behave as oxidizing as well reducing agents.

7. Statement 1: Oxidation number of carbon in HCN is +2.
Statement 2: Carbon always shows an oxidation state of +4.
Solution
(C) Carbon shows +2 oxidation state in HCN.
Here, the contribution of coordinate bond is neglected because the bond is directed from more electronegative nitrogen atom to less electronegative carbon atom.
Carbon shows different oxidation states in different compounds.

8. Statement 1: When $SnCl_2$ solution is added to $HgCl_2$ solution, a milky white precipitate is obtained and on adding excess $SnCl_2$, a black precipitate is formed.
Statement 2: The disproportionation of Hg(II) is easier than its reduction only.
Solution
(A)

9. Statement 1: Iodine shows oxidation state of +1 and +3 in the compounds ICl and ICl_3, respectively.
Statement 2: Iodine coming below the halogens F, Cl, Br in the halogen group of elements in the periodic table shows a higher degree of electropositive nature.
Solution
(A)

10. Statement 1: In aqueous solution, SO_2 reacts with H_2S liberating sulphur.
Statement 2: SO_2 is an effective reducing agent.
Solution
(C) SO_2 is an effective reducing as well as the oxidizing agent.

$$SO_2 + 2H_2S \rightarrow 3S + 2H_2O$$

11. Statement 1: In the titrations of Na_2CO_3 with HCl using methyl orange indicator, the volume required at the equivalence point is twice that of acid required using phenolphthalein as the indicator.
Statement 2: Two moles of HCl are required for complete neutralization of 1 mol of Na_2CO_3.
Solution
(B) $Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$ (phenolphthalein is the indicator)
$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$ (second end point with CH_3OH)
From the above reactions, it is clear that 2 mol of HCl are required for complete neutralization of Na_2CO_3. The titre value with methyl orange correspond to complete neutralization of Na_2CO_3 and with phenolphthalein correspond to half neutralization of Na_2CO_3.
12. Statement 1: The equivalent weight of reducing agent Fe^{2+} is 56 (atomic mass of Fe = 56).
Statement 2: Fe loses $2e^-$ to be converted into Fe^{2+}.
Solution

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(C) $Fe^{2+} \rightarrow Fe^{3+} + e^-$

13. Statement 1: Equivalent weight of a species can be written as molecular weight of the species divided by valence factor.

Statement 2: Valence factor represents valence in element, acidity in bases, basicity in acids, and total charge on cation or anion in an ionic compound

Solution

(B) Equivalent weight = Molecular weight/Valence factor

14. Statement 1: In solution A (200 mL of 0.1 N HCl) and solution B (100 mL of 0.2 N HCl), the equivalents of H^+ are the same.

Statement 2: Number of equivalents = Normality × Volume (in liters)

Solution

(A) Number of gram equivalent $= \frac{\text{Mass in grams}}{\text{Gram-equivalent mass}} = \frac{W(\text{g})}{\text{Gram-equivalent mass}}$

15. Statement 1: Addition of water to a solution containing solute and solvent changes its normality or molarity.

Statement 2: The milliequivalent and millimoles of the solute are not changed on dilution.

Solution

(B) $\text{Milliequivalent} = N \times V\,(\text{mL}) = \frac{\text{Weight}}{\text{Equivalent weight}} \times 1000.$

Comprehension Type

Read the paragraphs and answer the questions that follow.

Paragraph I

Aqueous solution of sodium hypochlorite (NaOCl) is household bleach and a strong oxidizing agent that reacts with chromite ion $[Cr(OH)_4]^-$ in basic solution to yield chromate (CrO_4^{2-}) and chloride ion.

1. Select the correct statement(s):

(A) OCl^- has been oxidized and $[Cr(OH)_4]^-$ has been reduced.

(B) OCl^- has been reduced and $[Cr(OH)_4]^-$ has been oxidized.

(C) It is simply a neutralization reaction.

(D) It is simply a displacement reaction.

Solution

(B) Loss of oxygen from OCl^- and loss of hydrogen from $[Cr(OH)_4]^-$

2. Balanced equation (only redox species) is

(A) $ClO^- + [Cr(OH)_4]^- \rightarrow CrO_4^{2-} + Cl^-$ (B) $3ClO^- + [Cr(OH)_4]^- \rightarrow CrO_4^{2-} + 3Cl^-$

(C) $3ClO^- + 2[Cr(OH)_4]^- \rightarrow 2CrO_4^{2-} + 3Cl^-$ (D) $ClO^- + 2[Cr(OH)_4]^- \rightarrow 2CrO_4^{2-} + Cl^-$

Solution

(C)

Paragraph II

For the unbalanced reaction $AX + BY + H_2O \rightarrow HA + OY + X_2B$. Let the oxidation number of X be –2

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and X, H_2O are not involved in redox reaction.

3. The elements undergoing oxidation is
(A) A. (B) B. (C) Y. (D) both B and Y.
Solution
(D) $\overset{+2}{A}+\overset{-2}{X}+\overset{+1}{B}Y+H_2O\rightarrow HA^{-1}+OY^{+2}+\overset{-2}{X_2}\overset{+4}{B}$ B, Y, or B or Y undergoes oxidation.

4. The possible oxidation states of B and Y in BY are, respectively,
(I) +1, –1 (II) +2, –2 (III) +3, –3
(A) I (B) II (C) III (D) I, II, III
Solution
(D) $\text{Oxidation states } B \leq +4; \text{oxidation states of } Y \leq 2$
Oxidation states of B = +1, Y = –1
Oxidation states of B = +2, Y = –2
Oxidation states of B = +3, Y = –3

5. If the above reaction is balanced in acid medium, the sum of smallest whole number stoichiometric coefficients of all the compounds will be
(A) 9. (B) 8. (C) 7. (D) 6.
Solution
(B) $2AX + BY + 2AX + BY + H_2O \rightarrow 2HA + OY + X_2B$

Paragraph III
Electron transfer or redox reactions involve simultaneous reduction and oxidation reactions. A substance is reduced when it gains electrons, and if it loses electrons it is said to be oxidized. A substance brings out reduction of other substances is called reducing agent, and the one which brings out oxidation of other substances is called oxidizing agent. In disproportionation, the same substance undergoes both oxidation and reduction. Two identical components of the same molecule or different molecules undergo oxidation and reduction to result a common product in this reaction called comproportionation reaction.

6. $2MnO_4^{2-}+2H_2O\rightarrow MnO_2+2MnO_4^-+4OH^-$ is an example of
(A) intermolecular redox reaction. (B) intramolecular redox reaction.
(C) disproportionation reaction. (D) comproportionation reaction.
Solution
(C) $NH_4NO_2 \rightarrow NH_4^+ NO_2^-;\ \overset{-3}{N} \rightarrow \overset{0}{N_2}$ and $\overset{+5}{N} \rightarrow N_2^0$

7. Which among the following is not a disproportionation reaction?
(A) $P_4 + OH^- \rightarrow H_2PO_4^- + PH_3$ (B) $S_2O_3^{2-} \rightarrow SO_4^{2-} + S$
(C) $H_2O_2 \rightarrow H_2O + O_2$ (D) $AgCl + NH_3 \rightarrow [Ag(NH_3)_2]Cl$
Solution
(D) $AgCl+NH_3 \rightarrow [Ag(NH_3)_2]Cl$ no change in oxidation state.

8. In the reaction $MnO_4^{2-}+2H_2O+2e^- \rightarrow MnO_2+4OH^-$, MnO_2 will act as

(A) an oxidizing agent. (B) reducing agent.
(C) both as an oxidizing and reducing agent. (D) precipitating agent.

Solution

(B) $\overset{+4}{\mathrm{Mn}}\mathrm{O_2} \rightarrow \overset{+6}{\mathrm{M}}\mathrm{nO_4^{2-}}$ reducing agent

Paragraph IV

A sample of iron(II) sulphate crystals, $FeSO_4 \cdot 7H_2O$ has been left open to the air and some of the iron(II) ions has been converted to iron(III). 4.2 g of the impure crystals were dissolved in a total of 250 cm^3 water and dilute sulphuric acid. 25 cm^3 portion of this solution was titrated with a solution of potassium bicarbonate(VI). The concentration of dichromate(VI) ions in this solution was 0.1 mol dm^{-3}. The average volume used was 23.5 cm^3.

9. How many moles of Fe^{2+} ions would there have been in the 250 cm^3 of stock solution?

(A) 7.05×10^{-4} mol (B) 2.35×10^{-4} mol (C) 1.41×10^{-2} (D) 7.05×10^{-4} mol

Solution

(C) 0.1 M $K_2Cr_2O_7$ was made in 23.5 cm^3 volume

$$\frac{0.1\,\text{mol}}{1000\,\text{mL}} \times 23.5 = 23.5\,\text{mL} = 2.35 \times 10^{-3}\ \text{mol of}\ \mathrm{Cr_2O_7^{2-}}$$

$$\mathrm{Cr_2O_7^{2-} + 14H^+ + 6}e^- \longrightarrow 2\mathrm{Cr^{3+}} + 7\mathrm{H_2O}$$

$$\mathrm{Fe^{2+} \longrightarrow Fe^{3+} + }e^- \times 6$$

$$\mathrm{Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \longrightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O}$$

$$\text{Therefore, number of moles of } \mathrm{Fe^{2+}} = 2.35 \times 6 \times 10^{-3}\ \text{mol} = 1.41 \times 10^{-3}\ \text{mol} / 25\,\text{cm}^3$$

$$\left.\begin{array}{l}\text{The number of moles of } \mathrm{Fe^{2+}} \text{ in} \\ 250\,\text{cm}^3 \text{ of solution}\end{array}\right\} \begin{array}{l}= 1.41 \times 10^{-3} \times 10\,\text{mol} \\ = 1.41 \times 10^{-2}\,\text{mol}\end{array}$$

10. What mass of Fe^{2+} ions should have been present in the 4.2 g of crystals?

(A) 0.84 g (B) 0.90 g (C) 0.77 g (D) 0.62 g

Solution

(A) Mass of Fe^{2+} = 0.015 mol × 56 = 0.84 g

11. The percentage purity of the crystal is

(A) 69% (B) 72% (C) 88% (D) 94%

Solution

(D) $\text{Number of moles of } \mathrm{FeSO_4 \cdot 7H_2O} = \text{Number of moles of } \mathrm{Fe^{2+}}$

$$\text{So, Number of moles of } \mathrm{FeSO_4 \cdot 7H_2O} = 1.41 \times 10^{-2}\ \text{mol}$$

$$\text{Grams of } \mathrm{FeSO_4 \cdot 7H_2O} = 1.41 \times 10^{-2}\ \text{mol} \times \text{Molecular mass of } \mathrm{FeSO_4 \cdot 7H_2O}$$

$$= 1.41 \times 10^{-2}\ \text{mol} \times 278 = 3.92\,\text{g}$$

$$\text{Percentage purity} = \frac{3.92\,\text{g}}{4.20\,\text{g}} \times 100 = 93.3\% \simeq 94\%$$

Paragraph V

1.00 g of a mixture having equal number of moles of carbonates of two alkali metals required 44.4 mL of 0.5 N HCl for complete reaction. Atomic weight of one of the metals is 7.00.

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12. The number of moles of each metal carbonate are

(A) 0.1 (B) 0.0111 (C) 0.0055 (D) 0.00275

Solution

(B) 1 mol of alkali metal carbonate would require 2 mol of HCl. As

$$M_2CO_3 + 2HCl \rightarrow MCl_2 + H_2O + CO_2$$

Hence, number of equivalents of each metal carbonates = 0.0111

So, Number of moles of metal carbonates = 0.0055

as *n*-factor = 2.

Hence, number of equivalents of HCl required to neutralize metal carbonates = 0.0222

Since alkali metal carbonates are equimolar, the number of equivalents of HCl required to neutralize each metal carbonates = 0.0222/2 = 0.0111.

13. The number of equivalents of HCl used is

(A) 0.222 (B) 2.22 (C) 22.22 (D) 0.0222

Solution

(D) Number of equivalents of acid is equal to $N \times V(\text{L}) = 0.5 \times \frac{44.4}{1000} = 0.0222$

Paragraph VI

Iodine titrations can be iodometric or iodimetric depending on using iodine directly or indirectly as an oxidizing agent in the redox titration.

(a) Iodimetric titrations in which a standard iodine solution is used as an oxidant and iodine is directly titrated against a reducing agent. For example,

$$2CuSO_4 + 4KI \rightarrow Cu_2I_2 + 2K_2SO_4 + I_2$$

(b) Iodimetric procedures are used for the determination of strength of reducing agent such as thiosulphates, sulphites, arsenites, and stannous chloride, by titrating them against standard solution of iodine taken in a burette.

$$2Na_2S_2O_3 + I_2 \rightarrow Na_2S_4O_6 + 2NaI$$

Starch is used as an indicator near the end point, which forms blue color complex with I_3^-. The blue color disappears when there is not more of free I_2.

14. In the reaction $2CuSO_4 + 4KI \rightarrow Cu_2I_2 + 2K_2SO_4 + I_2$, the equivalent weight of $CuSO_4$ is: (M_B: Molecular weight)

(A) $M_B/8$ (B) $M_B/4$ (C) $M_B/2$ (D) M_B

Solution

(D) $1e^- + Cu^{2+} \rightarrow Cu^+$

Therefore, Equivalent weight = $M_B/n = M_B$

15. The volume of KI solution used for $CuSO_4$ will be:

(A) 100 mL (B) 40 mL (C) 400 mL (D) 200 mL

Solution

(D) $\frac{15.95}{159.5} \times 1000 = 0.5 \times V \Rightarrow V = 200 \text{ mL}$

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16. When 319.0 g of $CuSO_4$ in a solution is reacted with excess of 0.5 M KI solution, then liberated iodine required 100 mL of 1.0 M $Na_2S_2O_3$ for complete reaction. The percentage purity of $CuSO_4$ in the sample is:

(A) 10% (B) 20% (C) 5% (D) None of these

Solution

(C) $CuSO_4 + KI \rightarrow I_2 \equiv Na_2S_2O_3$

Now, $\frac{w}{E} \times 1000 = 100 \times 1 \times 1 \Rightarrow w = 15.95$

So, $\%CuSO_4 = \frac{w}{319} \times 100 = 5\%$

Paragraph VII

Oleum is a mixture of H_2SO_4 and SO_3 (i.e., $H_2S_2O_7$). It is produced by passing SO_3 in H_2SO_4 solution. In order to dissolve free SO_3 in oleum, dilution of oleum is done, in which oleum converts into pure H_2SO_4. It is shown by the reaction as under:

$H_2SO_4 + SO_3 + H_2O \rightarrow 2H_2SO_4$ (Pure)

$SO_3 + H_2O \rightarrow H_2SO_4$ (Pure)

When 100 g sample of oleum is diluted with desired amount of H_2O (in grams), then the total mass of pure H_2SO_4 obtained after dilution is known as percentage labeling in oleum. Through this process, the percentage composition of H_2SO_4, SO_3 (free), and SO_3 (combined) is calculated.

If oleum sample is labeled as "109% H_2SO_4," that is, 100 g of oleum on dilution with 9 g of H_2O provides 109 g pure H_2SO_4, in which all free SO_3 in 100 g of oleum is dissolved.

17. For 109% labeled oleum, if the number of moles of H_2SO_4 and free SO_3 be x and y, respectively, then what will be the value of $x^2 + y^2$?

(A) 0.15 (B) 0.42 (C) 0.62 (D) 0.80

Solution

(C)

Therefore, 109% labeled oleum will contain 9 g H_2O, 40g free SO_3, 60 g H_2SO_4; as 109% means diluting 100 g of oleum gives 109 g of H_2SO_4.

40 g (free) SO_3 = Moles of SO_3 (free) = 0.5 = y

60 g (H_2SO_4) = 0.6122 mol H_2SO_4 = x

Therefore, $x^2 + y^2 = 0.25 + 0.37 = 0.62$

18. The percentage of combined SO_3 in the given oleum sample is

(A) 20% (B) 30% (C) 48.98% (D) 51 %

Solution

(C) 98 g H_2SO_4 contains 80 g SO_3

60 g H_2SO_4 contains $= \frac{80 \times 60}{98} = 48.98$ g

% of combined SO_3 in oleum = 48.98%

19. The percentage composition of free SO_3 and H_2SO_4 in the oleum sample, respectively, are

(A) 60%, 40% (B) 30%, 70% (C) 85%, 15% (D) 40%, 60%.

Solution

(D) In 100 g oleum sample,

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$$\underset{0.5\text{ mol}}{SO_3} + H_2O \rightarrow H_2SO_4$$

Therefore, SO_3 = 0.5 mol = 40 g

Weight of SO_3 = 40 g

Weight of H_2SO_4 = 60 g

% of SO_3 = 40%

% of H_2SO_4 = 60%

20. What volume of 1 M NaOH (in milliliters) will be required to react completely with H_2SO_4 and SO_3?

(A) 250 mL (B) 2224 mL (C) 750 mL (D) 1800 mL.

Solution

(B) As milliequivalent of NaOH = Milliequivalent of SO_3 + Milliequivalent of H_2SO_4

Therefore, $$1\times1\times V = \frac{40}{80}\times2 + \frac{60}{98}\times2 \Rightarrow V = 2.224\text{ L}$$

Integer Answer Type

The answer is **non-negative integer**.

1. The number of electrons involved in the conversion of MnO_4^- to MnO_2 is ____.

Solution

(3) $MnO_4^- + 2H_2O + 3e^- \quad MnO_2 + 4OH^-$

2. In $Ba(H_2PO_2)_2$ the oxidation number of phosphorus is ____.

Solution

(1) $(H_2PO_2)_2^{2-}$

$2(2(+1)) + x + 2(-2) = -2 \Rightarrow x = +1$

3. In bleaching powder ($CaOCl_2$), the oxidation states of chlorine are *x* and *y*. Value of (redox can be removed) *x* + *y* is ____.

Solution

(0) As the molecule is neutral.

4. Given balanced chemical equation for oxidation of phosphorus (III) sulphide by nitric acid. The products include NO and SO_2.

$$aP_4S_6 + bH^+ + cNO_3^- \rightarrow dNO + eH_3PO_4 + fSO_2 + gH_2O$$

What is the value of (*a* + *g*)?

Solution

(7) The balanced redox reaction is

$$3P_4S_6 + 44H^+ + 44NO_3^- \rightarrow 44NO + 12H_3PO_4 + 18SO_2 + 4H_2O$$

5. In the redox reaction, $xNO_3^- + yAs_2S_3 + zH_2O \rightarrow AsO_4^{3-} + NO^+ + SO_4^{2-} + H^+$ what is the value of (*x*/*y*)?

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Solution
(14)

The half reactions are

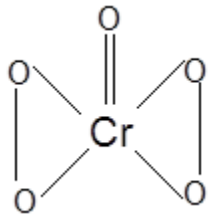
$$2\overset{+5}{\mathrm{N}}\mathrm{O}_3^- + 8\mathrm{H}^+ + 2e^- \rightarrow 2\overset{+3}{\mathrm{N}}\mathrm{O}^+ + 4\mathrm{H}_2\mathrm{O}] \times 5$$

$$\overset{+3}{\mathrm{As}}_2\overset{-2}{\mathrm{S}}_3 + 20\mathrm{H}_2\mathrm{O} \rightarrow 2\overset{+5}{\mathrm{As}}\mathrm{O}_4^{3-} + 3\overset{+6}{\mathrm{S}}\mathrm{O}_4^{2-} + 10e^- + 40\mathrm{H}^+$$

Overall balanced reaction

$$14\overset{+5}{\mathrm{N}}\mathrm{O}_3^- + \overset{+3}{\mathrm{As}}_2\overset{-2}{\mathrm{S}}_3 + 16\mathrm{H}^+ \rightarrow 14\overset{+3}{\mathrm{N}}\mathrm{O}^+ + 2\overset{+5}{\mathrm{As}}\mathrm{O}_4^{3-} + 3\overset{+6}{\mathrm{S}}\mathrm{O}_4^{2-} + 8\mathrm{H}_2\mathrm{O}$$

6. Oxidation state of Cr in CrO_5 is ___.
Solution
(7) The structure is



7. 0.01 mol of FeS_n (iron (II) sulphide) required 0.06 mol of AO_4^{3-} for complete oxidation. The species formed are FeO, SO_2, and A^{2+}. The value of *n* is ___.
Solution
(4) 0.01 × *n*-factor = 0.06 × 3 (as Milliequivalent of FeS_n = Milliequivalent of AO_4^{3-})
As *n*-factor = 18 (for FeS_n), we have

$$18 = \left(\frac{2}{n} + 4\right) \times n \Rightarrow n = 4$$

8. 1575 g of oxalic acid $(COOH)_2 \cdot xH_2O$ are dissolved in water and the volume made upto 250 mL. On titration, 16.68 mL of this solution requires 25mL of N/15 NaOH solution for complete neutralization. The value of *x* is ___.
Solution

Milliequivalent of oxalic acid in 16.68 mL = Milliequivalent of NaOH $= 25 \times \frac{1}{15}$

Milliequivalent of oxalic acid in 250 mL $= 25 \times \frac{1}{15} \times \frac{250}{16.68} = 24.98$

Therefore, $\frac{1.575}{(90 + 18x)/2} \times 1000 = 24.98$

n-Factor of oxalic acid is 2 as dibasic acid. Solving *x* = 2, that is, $(COOH)_2 \cdot 2H_2O$
9. The equivalent weight of a metal carbonate 0.84 g of which reacts exactly with 40 mL of N/2 H_2SO_4 is ____.

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Solution

(42) $\frac{0.84}{E} = \frac{40 \times 1}{2 \times 1000} \Rightarrow E = 42$

10. 5 g of K_2SO_4 were dissolved in 250 mL of solution. The volume of this solution that should be used so that 1.2 g of $BaSO_4$ may be precipitated from $BaCl_2$ is ___. (Given that molar mass of K_2SO_4 = 174 and $BaSO_4$ = 233.)

Solution

(45 mL) $BaCl_2 + K_2SO_4 \rightarrow BaSO_4 \downarrow + 2KCl$

Therefore, $n_{K_2SO_4} = n_{BaSO_4}$

Matrix–Match Type

1. Match the compound with the average oxidation state of Fe.

Column I	**Column II**
(A) $Fe_4[Fe(CN)_6]_3$	(p) 12/5
(B) $[Fe(NO)(H_2O)_5]SO_4$	(q) 2
(C) $Fe_3[Fe(CN)_6]_2$	(r) 18/7
(D) $Na_2[Fe(CN)_5NO]$	(s) 1

Solution

(A) → (r); (B) → (s); (C) → (p); (D) → (q)

(A) $7x + (-18) = 0$ (B) $x + 1 - 2 = 0$

(C) $5x - 12 = 0$ (D) $+2 + x - 5 + 1 = 0$

2. Match the reaction with its type.

Column I	**Column II**
(A) $V_2O_5Ca \rightarrow 2V + 5CaO$	(p) Non-metal displacement reaction
(B) $2NaH \rightarrow 2Na + H_2$	(q) Disproportionation reaction
(C) $P_4 + 3OH^- + 3H_2O \rightarrow PH_3 + 3H_2PO_2$	(r) Decomposition reaction
(D) $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$	(s) Redox reaction

Solution

(A) → (s); (B) → (r, s); (C) → (q, s); (D) → (p, s)

(A) $\overset{+5}{V_2}O_5 + 5\overset{0}{Ca} \rightarrow 2\overset{0}{V} + 5\overset{+2}{Ca}O$

This is a redox reaction.

(B) $2\overset{+1}{Na}\overset{-1}{H} \rightarrow 2\overset{0}{Na} + \overset{0}{H_2}$

This is an example of decomposition as well as redox reaction.

(C) $\overset{0}{P_4} + 3\overset{-}{OH} + 3H_2O \rightarrow \overset{-3}{P}H_3 + 3H_2\overset{+2}{P}O_2$

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Disproportionation is a reaction in which the same element undergoes oxidation as well as reduction. It is also a type of redox reaction.

(D) $\overset{0}{Ca} + 2\overset{+1}{H}_2\overset{-2}{O} \rightarrow \overset{+2}{Ca}(OH)_2 + \overset{0}{H}_2$

This reaction is an example of non-metal displacement reaction. It is also a redox reaction.

3. Match the characteristic properties with the elements.

Column I	**Column II**
(A) Metal which reacts with dilute acids to give H_2 gas.	(p) Zinc
(B) Metal whose container can be used to store conc. HNO_3.	(q) Silver
(C) Metal which is used as an electrode in Daniell cell.	(r) Aluminium
(D) Metal which does not react with dilute acids to give H_2 gas.	(s) Copper

Solution

(A) → (p, r); (B) → (r); (C) → (p, s); (D) → (q, s)

(A) $Zn + 2HCl \rightarrow H_2 \uparrow + ZnCl_2$

$2Al + 6HCl \rightarrow 3H_2 \uparrow + 2AlCl_3$

(B) Aluminium container can be used to store conc. HNO_3 because Al is rendered passive by nitric acid. This is due to oxidation and formation of a thin film of oxide on its surface. Hence Al does not react with conc. HNO_3 so it can be used to store HNO_3.

(C) Zn and Cu can be used as electrode in Daniel cell

$Zn \rightarrow Zn^{2+} + 2e^-$ (Oxidation)

$Cu^{2+} + 2e^- \rightarrow Cu$ (Reduction)

(D) Silver and copper do not react with dilute acids to give H_2 gas because it acts as oxidizing agent and it has higher reduction potential value.

4. Match the bold atoms in compounds with their oxidation numbers.

Column I	**Column II**
(A) $\mathbf{H}_2O_2$	(p) +1
(B) Mn$\mathbf{S}O_4$	(q) +3
(C) $\mathbf{Al}Cl_3$	(r) +2
(D) $\mathbf{S}_2\mathbf{Cl}_2$	(s) +6

Solution

(A) → (r); (B) → (s); (C) → (q); (D) → (p, r)

(A) $2x + 2(-2) = 0$ so $x = +2$

(B) $x + (+2) + (-2) \times 4 = 0$ $x = +6$

(C) $x + (-1) \times 3 = 0$ so $x = +3$

(D) $2x + (-1)2 = 0$ so $x = +1$ and $2(-2) + 2y = 0$ so $y = +2$

5. Match the reactions with their equivalent weights of their oxidizing/reducing agents.

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Column I	Column II
(A) $NH_3 \rightarrow \overset{-}{N}O_3$	(p) $M/3$
(B) $FeC_2O_4 \rightarrow Fe^{3+} + 2CO_3^{2-}$	(q) $M/6$
(C) $H_2SO_5 \rightarrow S_8$	(r) $M/8$
(D) $KMnO_4 \rightarrow Mn^{2+}$	(s) $M/5$
	(t) Reducing agent

Solution

A → (r, t), B → (p, t), C → (q), D → (s)

$\overset{-3}{N}H_3 \rightarrow \overset{+5}{N}O_3^-$ *n*-factor = 8, Equivalent weight = Molecular weight/(+5 – (–3)) = Molecular weight/8

$\overset{+2}{Fe}\overset{+3}{C_2}O_4\overset{+3}{Fe} + 2\overset{+4}{C}O_3^{2-}$ *n*-factor = 3, Equivalent weight = Molecular weight/3

$H_2\overset{+6}{S}O_5 \rightarrow \overset{0}{S}_8$ *n*-factor = 6, Equivalent weight = Molecular weight/6

$K\overset{+7}{M}nO_4 \rightarrow \overset{+2}{M}n$ *n*-factor = 5, Equivalent weight = Molecular weight/5

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