

NCERT EXERCISES

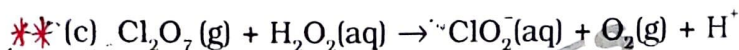
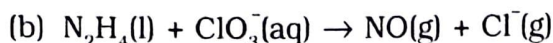
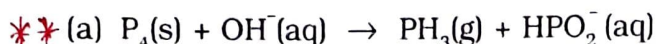
- 8.1 Assign oxidation number to the underlined elements in each of the following species:
 (a) $\text{NaH}_2\underline{\text{P}}\text{O}_4$ (b) $\text{NaH}\underline{\text{S}}\text{O}_4$ (c) $\text{H}_4\underline{\text{P}}_2\text{O}_7$ (d) $\text{K}_2\underline{\text{Mn}}\text{O}_4$
 (e) $\text{Ca}\underline{\text{O}}_2$ (f) $\text{Na}\underline{\text{B}}\text{H}_4$ * (g) $\text{H}_2\underline{\text{S}}_2\text{O}_7$ * (h) $\text{KAl}(\underline{\text{S}}\text{O}_4)_2 \cdot 12 \text{H}_2\text{O}$
- 8.2 What are the oxidation number of the underlined elements in each of the following and how do you rationalise your results ?
 (a) $\text{K}\underline{\text{I}}_3$ * (b) $\text{H}_2\underline{\text{S}}_4\text{O}_6$ * (c) $\underline{\text{F}}\text{e}_3\underline{\text{O}}_4$ * (d) $\underline{\text{C}}\text{H}_3\underline{\text{C}}\text{H}_2\underline{\text{O}}\text{H}$ * (e) $\underline{\text{C}}\text{H}_3\underline{\text{C}}\text{O}\underline{\text{O}}\text{H}$
- 8.3 Justify that the following reactions are redox reactions:
 (a) $\text{CuO}(\text{s}) + \text{H}_2(\text{g}) \rightarrow \text{Cu}(\text{s}) + \text{H}_2\text{O}(\text{g})$
 (b) $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + 3\text{CO}_2(\text{g})$
 (c) $4\text{BCl}_3(\text{g}) + 3\text{LiAlH}_4(\text{s}) \rightarrow 2\text{B}_2\text{H}_6(\text{g}) + 3\text{LiCl}(\text{s}) + 3\text{AlCl}_3(\text{s})$
 (d) $2\text{K}(\text{s}) + \text{F}_2(\text{g}) \rightarrow 2\text{K}^+\text{F}^-(\text{s})$
 (e) $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g})$
- 8.4 Fluorine reacts with ice and results in the change:
 $\text{H}_2\text{O}(\text{s}) + \text{F}_2(\text{g}) \rightarrow \text{HF}(\text{g}) + \text{HOF}(\text{g})$
 Justify that this reaction is a redox reaction.
- 8.5 Calculate the oxidation number of sulphur, chromium and nitrogen in H_2SO_5 , $\text{Cr}_2\text{O}_7^{2-}$ and NO_3^- . Suggest structure of these compounds. Count for the fallacy.
- 8.6 Write formulas for the following compounds:
 (a) Mercury(II) chloride (b) Nickel(II) sulphate
 (c) Tin(IV) oxide (d) Thallium(I) sulphate
 (e) Iron(III) sulphate (f) Chromium(III) oxide
- 8.7 Suggest a list of the substances where carbon can exhibit oxidation states from -4 to +4 and nitrogen from -3 to +5.
- 8.8 While sulphur dioxide and hydrogen peroxide can act as oxidising as well as reducing agents in their reactions, ozone and nitric acid act only as oxidants. Why ?
- 8.9 Consider the reactions:
 (a) $6\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(\text{aq}) + 6\text{O}_2(\text{g})$
 (b) $\text{O}_3(\text{g}) + \text{H}_2\text{O}_2(\text{l}) \rightarrow \text{H}_2\text{O}(\text{l}) + 2\text{O}_2(\text{g})$
 Why it is more appropriate to write these reactions as :
 (a) $6\text{CO}_2(\text{g}) + 12\text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(\text{aq}) + 6\text{H}_2\text{O}(\text{l}) + 6\text{O}_2(\text{g})$
 (b) $\text{O}_3(\text{g}) + \text{H}_2\text{O}_2(\text{l}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g}) + \text{O}_2(\text{g})$
 Also suggest a technique to investigate the path of the above (a) and (b) redox reactions.
- 8.10 The compound AgF_2 is unstable compound. However, if formed, the compound acts as a very strong oxidising agent. Why ?
- 8.11 Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. Justify this statement giving three illustrations.

- 8.12 How do you count for the following observations ?
- (a) Though alkaline potassium permanganate and acidic potassium permanganate both are used as oxidants, yet in the manufacture of benzoic acid from toluene we use alcoholic potassium permanganate as an oxidant. Why ? Write a balanced redox equation for the reaction.
- (b) When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smelling gas HCl, but if the mixture contains bromide then we get red vapour of bromine. Why ?
- 8.13 Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions:
- (a) $2\text{AgBr (s)} + \text{C}_6\text{H}_6\text{O}_2\text{(aq)} \rightarrow 2\text{Ag(s)} + 2\text{HBr (aq)} + \text{C}_6\text{H}_4\text{O}_2\text{(aq)}$
- (b) $\text{HCHO(l)} + 2[\text{Ag (NH}_3)_2]^+\text{(aq)} + 3\text{OH}^-\text{(aq)} \rightarrow 2\text{Ag(s)} + \text{HCOO}^-\text{(aq)} + 4\text{NH}_3\text{(aq)} + 2\text{H}_2\text{O(l)}$
- (c) $\text{HCHO (l)} + 2 \text{Cu}^{2+}\text{(aq)} + 5 \text{OH}^-\text{(aq)} \rightarrow \text{Cu}_2\text{O(s)} + \text{HCOO}^-\text{(aq)} + 3\text{H}_2\text{O(l)}$
- (d) $\text{N}_2\text{H}_4\text{(l)} + 2\text{H}_2\text{O}_2\text{(l)} \rightarrow \text{N}_2\text{(g)} + 4\text{H}_2\text{O(l)}$
- (e) $\text{Pb(s)} + \text{PbO}_2\text{(s)} + 2\text{H}_2\text{SO}_4\text{(aq)} \rightarrow 2\text{PbSO}_4\text{(s)} + 2\text{H}_2\text{O(l)}$
- 8.14 Consider the reactions :
- $2 \text{S}_2\text{O}_3^{2-}\text{(aq)} + \text{I}_2\text{(s)} \rightarrow \text{S}_4\text{O}_6^{2-}\text{(aq)} + 2\text{I}^-\text{(aq)}$
- $\text{S}_2\text{O}_3^{2-}\text{(aq)} + 2\text{Br}_2\text{(l)} + 5 \text{H}_2\text{O(l)} \rightarrow 2\text{SO}_4^{2-}\text{(aq)} + 4\text{Br}^-\text{(aq)} + 10\text{H}^+\text{(aq)}$
- Why does the same reductant, thiosulphate react differently with iodine and bromine ?
- 8.15 Justify giving reactions that among halogens, fluorine is the best oxidant and among hydrohalic compounds, hydroiodic acid is the best reductant.
- 8.16 Why does the following reaction occur ?
- $\text{XeO}_6^{4-}\text{(aq)} + 2\text{F}^-\text{(aq)} + 6\text{H}^+\text{(aq)} \rightarrow \text{XeO}_3\text{(g)} + \text{F}_2\text{(g)} + 3\text{H}_2\text{O(l)}$
- What conclusion about the compound Na_4XeO_6 (of which XeO_6^{4-} is a part) can be drawn from the reaction.
- 8.17 Consider the reactions:
- (a) $\text{H}_3\text{PO}_2\text{(aq)} + 4 \text{AgNO}_3\text{(aq)} + 2 \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{PO}_4\text{(aq)} + 4\text{Ag(s)} + 4\text{HNO}_3\text{(aq)}$
- (b) $\text{H}_3\text{PO}_2\text{(aq)} + 2\text{CuSO}_4\text{(aq)} + 2 \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{PO}_4\text{(aq)} + 2\text{Cu(s)} + \text{H}_2\text{SO}_4\text{(aq)}$
- (c) $\text{C}_6\text{H}_5\text{CHO(l)} + 2[\text{Ag (NH}_3)_2]^+\text{(aq)} + 3\text{OH}^-\text{(aq)} \rightarrow \text{C}_6\text{H}_5\text{COO}^-\text{(aq)} + 2\text{Ag(s)} + 4\text{NH}_3\text{(aq)} + 2 \text{H}_2\text{O(l)}$
- (d) $\text{C}_6\text{H}_5\text{CHO(l)} + 2\text{Cu}^{2+}\text{(aq)} + 5\text{OH}^-\text{(aq)} \rightarrow$ No change observed.
- What inference do you draw about the behaviour of Ag^+ and Cu^{2+} from these reactions ?
- 8.18 Balance the following redox reactions by ion - electron method :
- * (a) $\text{MnO}_4^-\text{(aq)} + \text{I}^-\text{(aq)} \rightarrow \text{MnO}_2\text{(s)} + \text{I}_2\text{(s)}$ (in basic medium)
- (b) $\text{MnO}_4^-\text{(aq)} + \text{SO}_2\text{(g)} \rightarrow \text{Mn}^{2+}\text{(aq)} + \text{HSO}_4^-\text{(aq)}$ (in acidic solution)
- (c) $\text{H}_2\text{O}_2\text{(aq)} + \text{Fe}^{2+}\text{(aq)} \rightarrow \text{Fe}^{3+}\text{(aq)} + \text{H}_2\text{O(l)}$ (in acidic solution)
- ** (d) $\text{Cr}_2\text{O}_7^{2-} + \text{SO}_2\text{(g)} \rightarrow \text{Cr}^{3+}\text{(aq)} + \text{SO}_4^{2-}\text{(aq)}$ (in acidic solution)

REDOX REACTIONS

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

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



8.21 The Mn^{3+} ion is unstable in solution and undergoes disproportionation to give Mn^{2+} , MnO_2 , and H^+ ion. Write a balanced ionic equation for the reaction.

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8.22 Consider the elements :

Cs, Ne, I and F

(a) Identify the element that exhibits only negative oxidation state.

(b) Identify the element that exhibits only positive oxidation state.

(c) Identify the element that exhibits both positive and negative oxidation states.

(d) Identify the element which exhibits neither the negative nor does the positive oxidation state.

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

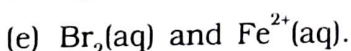
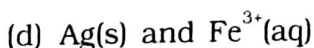
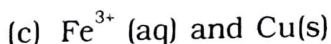
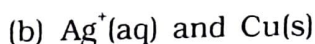
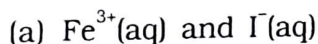
8.24 Refer to the periodic table given in your book and now answer the following questions:

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

(b) Select three metals that can show disproportionation reaction.

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

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



8.27 Predict the products of electrolysis in each of the following:

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- (i) An aqueous solution of AgNO_3 with silver electrodes
- (ii) An aqueous solution AgNO_3 with platinum electrodes
- (iii) A dilute solution of H_2SO_4 with platinum electrodes
- (iv) An aqueous solution of CuCl_2 with platinum electrodes.

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

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Al, Cu, Fe, Mg and Zn.

8.29 Given the standard electrode potentials,

* *

$\text{K}^+/\text{K} = -2.93\text{V}$, $\text{Ag}^+/\text{Ag} = 0.80\text{V}$,

$\text{Hg}^{2+}/\text{Hg} = 0.79\text{V}$

$\text{Mg}^{2+}/\text{Mg} = -2.37\text{V}$, $\text{Cr}^{3+}/\text{Cr} = -0.74\text{V}$

arrange these metals in their increasing order of reducing power.

8.30 Depict the galvanic cell in which the reaction $\text{Zn(s)} + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag(s)}$ takes place. Further show:

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- (i) which of the electrode is negatively charged,
- (ii) the carriers of the current in the cell, and
- (iii) individual reaction at each electrode.

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