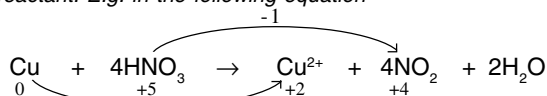


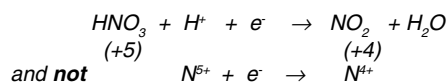




**Exam Hint:** - When describing redox reactions some candidates use the oxidation state of the element rather than the **actual** formula of the reactant. E.g. in the following equation



$\text{HNO}_3$  is the oxidising agent, **not**  $\text{N}^{5+}$  - There is no such ion!  
The redox half-equation is



### Questions

- Which of the following equations are redox reactions?
  - $\text{Br}_2 + 2\text{I}^- \rightarrow \text{I}_2 + 2\text{Br}^-$
  - $3\text{CuO} + 2\text{NH}_3 \rightarrow 3\text{Cu} + \text{N}_2 + 3\text{H}_2\text{O}$
  - $2\text{Pb}(\text{NO}_3)_2 \rightarrow 2\text{PbO} + 4\text{NO}_2 + \text{O}_2$
  - $2[\text{Fe}(\text{H}_2\text{O})_6]^{3+} + 3\text{CO}_3^{2-} \rightarrow 2[\text{Fe}(\text{OH})_3(\text{H}_2\text{O})_3] + 3\text{CO}_2 + 3\text{H}_2\text{O}$
  - $\text{Na}_2\text{S}_2\text{O}_8 + 2\text{NaI} \rightarrow \text{I}_2 + 2\text{Na}_2\text{SO}_4$
  - $2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2$
  - $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$
  - $\text{Mg} + \text{Zn}^{2+} \rightarrow \text{Mg}^{2+} + \text{Zn}$
  - $\text{K}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{KCl} + \text{H}_2\text{O} + \text{CO}_2$
  - $\text{Ag}^+(\text{aq}) + \text{I}^-(\text{aq}) \rightarrow \text{AgI}(\text{s})$
- Identify the *oxidising agent* in each of the following equations:
    - $3\text{Cl}_2 + 2\text{Fe} \rightarrow 2\text{FeCl}_3$
    - $2\text{H}_2\text{S} + \text{SO}_2 \rightarrow 3\text{S} + 2\text{H}_2\text{O}$
    - $\text{Cu} + 4\text{HNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + 2\text{H}_2\text{O} + 2\text{NO}_2$
    - $5\text{H}_2\text{O}_2 + 2\text{MnO}_4^- + 6\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 5\text{O}_2 + 8\text{H}_2\text{O}$
    - $\text{H}_2\text{O}_2 + \text{H}_2\text{S} \rightarrow \text{S} + 2\text{H}_2\text{O}$
  - Identify the *reducing agent* in each of the following equations:
    - $\text{Ca} + 2\text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2 + \text{H}_2$
    - $\text{Cl}_2 + 2\text{I}^- \rightarrow 2\text{Cl}^- + \text{I}_2$
    - $\text{Cu} + 2\text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{SO}_2 + 2\text{H}_2\text{O}$
    - $3\text{SO}_3^{2-} + \text{Cr}_2\text{O}_7^{2-} + 8\text{H}^+ \rightarrow 3\text{SO}_4^{2-} + 2\text{Cr}^{3+} + 4\text{H}_2\text{O}$
    - $2\text{VO}^{2+} + \text{Zn} + 4\text{H}^+ \rightarrow 2\text{V}^{3+} + \text{Zn}^{2+} + 2\text{H}_2\text{O}$
- $\text{Fe}^{2+}$  is changed into  $\text{Fe}^{3+}$  by reaction with acidified  $\text{Cr}_2\text{O}_7^{2-}$  which becomes  $\text{Cr}^{3+}$ . Using redox  $\frac{1}{2}$ -equations construct the overall redox equation for this reaction. Which species is the oxidising agent? Give a reason for your answer.
  - During the course of a redox reaction between iodine,  $\text{I}_2$  and the thiosulphate ion,  $\text{S}_2\text{O}_3^{2-}$ , the products formed are the iodide,  $\text{I}^-$  and tetrathionate,  $\text{S}_4\text{O}_6^{2-}$  ions. Using redox  $\frac{1}{2}$ -equations, construct the overall equation for the reaction. Identify the reducing agent, giving reasons for your decision.
- From the two equations given below, construct and identify the two redox  $\frac{1}{2}$ -equations:
  - $2\text{I}^- + \text{H}_2\text{O}_2 + 2\text{H}^+ \rightarrow \text{I}_2 + 2\text{H}_2\text{O}$
  - $\text{I}^- + \text{IO}_3^- + 6\text{H}^+ \rightarrow \text{I}_2 + 3\text{H}_2\text{O}$
- The manganate(VI) ion in acid solution undergoes a disproportionation reaction as follows:
 
$$3\text{MnO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) \rightarrow 2\text{MnO}_4^- + \text{MnO}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l})$$
  - What is a disproportionation reaction?
  - Show, using both oxidation numbers and  $\frac{1}{2}$ -equations, that this is an example of such a reaction.

### Answers

- $\text{Cl}$  changes oxidation number from 0 to  $-1$ ,  $\text{Br}$  from  $-1$  to 0).
  - $\text{Cu}$  from  $+2$  to 0,  $\text{N}$  from  $-3$  to 0).
  - $\text{N}$  from  $+5$  to  $+4$ ,  $\text{O}$  from  $-2$  to 0 in  $\text{O}_2$
  - $\text{S}$  from  $+7$  to  $+6$ ,  $\text{I}$  from  $-1$  to 0).
  - $\text{Na}$  from 0 to  $+1$ ,  $\text{Cl}$  from 0 to  $-1$
  - $\text{Mg}$  from 0 to  $+2$ ,  $\text{Zn}$  from  $2+$  to 0)
- $\text{Cl}_2$ : ( $\text{Cl}$  decreases oxidation number from 0 to  $-1$  and is reduced).
    - $\text{SO}_2$  ( $\text{S}$  from  $+4$  to 0)
    - $\text{HNO}_3$  ( $\text{N}$  from  $+5$  to  $+4$  in  $\text{NO}_2$ ).
    - $\text{MnO}_4^-$  ( $\text{Mn}$  from  $+7$  to  $+2$ ).
    - $\text{H}_2\text{O}_2$  ( $\text{O}$  from  $-1$  to  $-2$ ). [Note that the oxidation number of  $\text{O}$  in  $\text{H}_2\text{O}_2$  is  $-1$ , not  $-2$ . If it was  $-2$ ,  $\text{H}$  would be  $+2$  which is not possible because a  $\text{H}$  atom only has one electron and cannot form  $\text{H}^{2+}$ ].
  - $\text{Ca}$  (increases oxidation number from 0 to  $+2$  and is oxidised).
    - $\text{I}^-$  ( $\text{I}$  from  $-1$  to 0).
    - $\text{Cu}$  (from 0 to  $+2$ ).
    - $\text{SO}_3^{2-}$  ( $\text{S}$  from  $+4$  to  $+6$ ).
    - $\text{Zn}$  (from 0 to  $+2$ ).
- $$[\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-] \times 6 \quad \text{OXIDATION}$$

$$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} \quad \text{REDUCTION}$$

$$6\text{Fe}^{2+} + \text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ \rightarrow 6\text{Fe}^{3+} + 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$$

$$\text{Cr}_2\text{O}_7^{2-} \text{ (acidified) is the oxidising agent. It receives electrons from } \text{Fe}^{2+} \text{ and is reduced.}$$
  - $$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^- \quad \text{REDUCTION}$$

$$2\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4\text{O}_6^{2-} + 2\text{e}^- \quad \text{OXIDATION}$$

$$\text{I}_2 + 2\text{S}_2\text{O}_3^{2-} \rightarrow 2\text{I}^- + \text{S}_4\text{O}_6^{2-}$$

$$\text{S}_2\text{O}_3^{2-} \text{ is the reducing agent. It gives electrons to } \text{I}_2 \text{ and is oxidised.}$$
- $$2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^- \quad \text{OXIDATION}$$

$$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O} \quad \text{REDUCTION}$$
  - $$2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^- \quad \text{OXIDATION}$$

$$2\text{IO}_3^- + 12\text{H}^+ + 10\text{e}^- \rightarrow \text{I}_2 + 6\text{H}_2\text{O}$$
- A disproportionation reaction is one in which the same element is both oxidised and reduced.
  - Manganese is both oxidised and reduced:
 
$$\text{MnO}_4^{2-} \rightarrow \text{MnO}_4^- + \text{e}^- \quad \text{OXIDATION}$$

$$+6 \quad \quad \quad +7$$

$$\text{MnO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{MnO}_2 + 2\text{H}_2\text{O} \quad \text{REDUCTION}$$

$$+6 \quad \quad \quad \quad \quad +4$$

**Acknowledgements:** This Factsheet was researched and written by Derek Swain. Curriculum Press, Bank House, 105 King Street, Wellington, Shropshire, TF1 1NU. ChemistryFactsheets may be copied free of charge by teaching staff or students, provided that their school is a registered subscriber. No part of these Factsheets may be reproduced, stored in a retrieval system, or transmitted, in any other form or by any other means, without the prior permission of the publisher. ISSN 1351-5136