

Exercise walk-through

6.1

Redox processes: electron transfer and changes in oxidation number (oxidation state)

A-Level Chemistry

Oxidation number

The **oxidation number** shows electrons lost or gained vs the free element.

- element = 0; simple ion = its charge
- H = +1, O = -2, F = -1
- they sum to 0 (or the ion's charge)

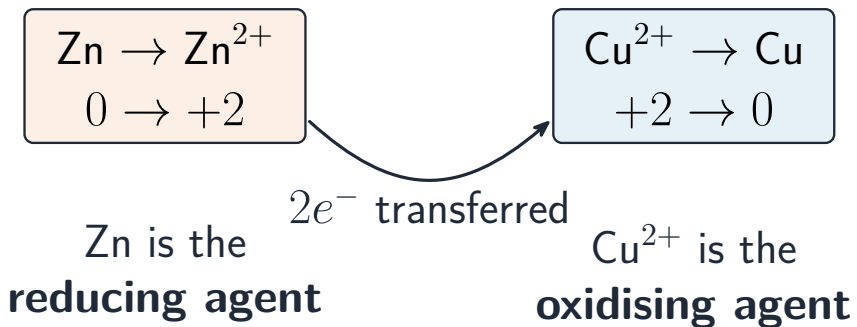


they sum to 0: $(+1) + \text{Mn} + (-8) = 0$, so $\text{Mn} = +7$

Redox and OIL RIG

oxidation: loses $2e^-$
(oxidation number $\uparrow$)

reduction: gains $2e^-$
(oxidation number $\downarrow$)

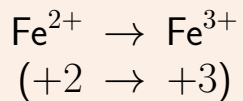


A **redox** reaction transfers electrons:

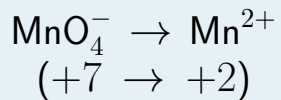
- **oxidation** is loss
(number up)
- **reduction** is gain
(number down)

Remember **OIL RIG**; they always happen together.

Balancing redox equations



oxidised: rises 1 ($\times 5$)



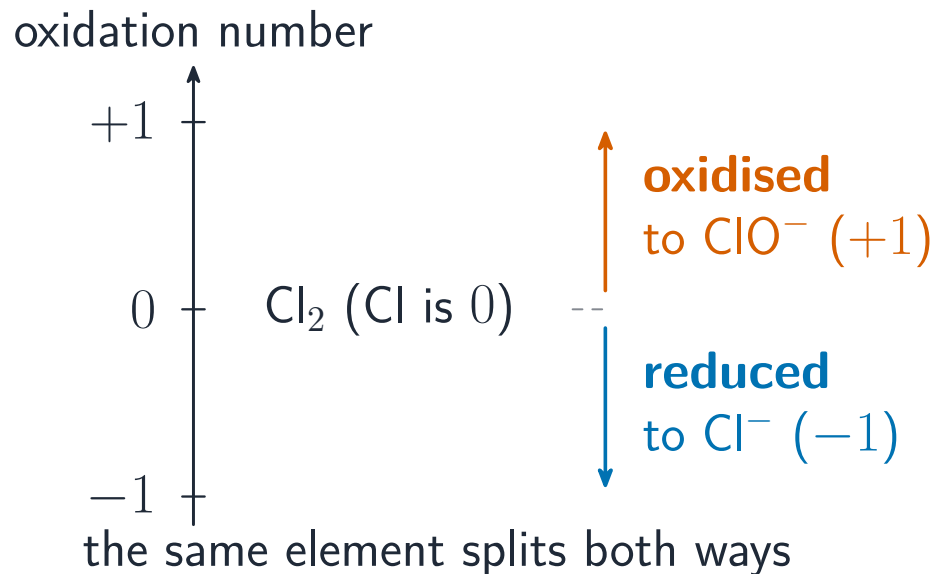
reduced: falls 5 ($\times 1$)

5 electrons lost = 5 gained $\Rightarrow$ ratio **5 Fe²⁺ : 1 MnO₄⁻**

Use oxidation-number changes:

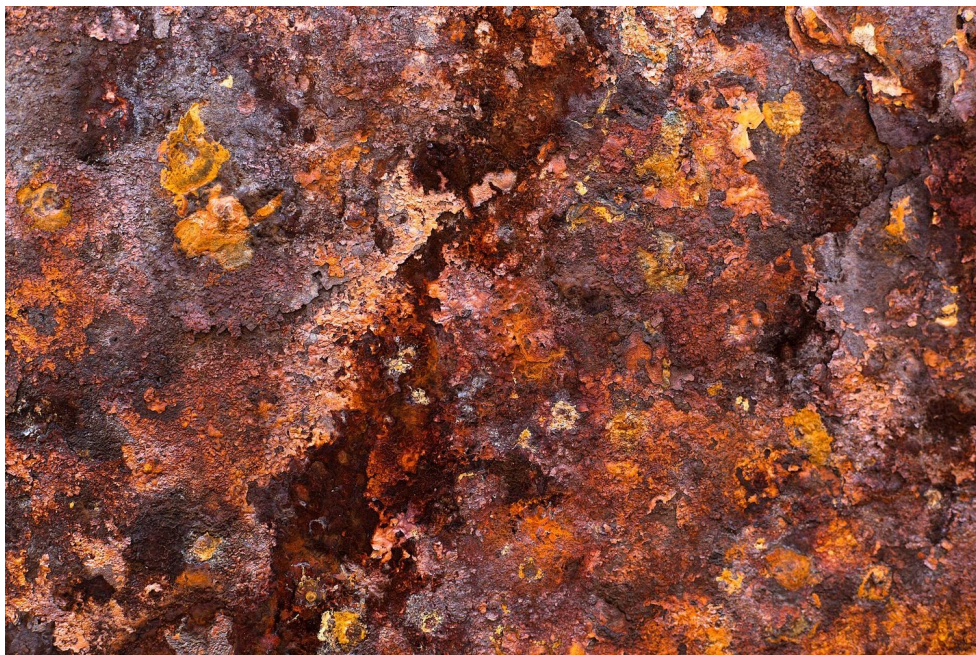
The total **rise** in oxidation number must equal the total **fall** (every electron lost is gained) – this fixes the ratio.

Disproportionation



Disproportionation is when the **same** element is both oxidised and reduced at once. Chlorine (0) $\rightarrow$ ClO^- (+1) *and* Cl^- (-1) in cold water.

Oxidising and reducing agents



- an **oxidising agent**
takes electrons – it is itself
reduced
- a **reducing agent**
gives electrons – it is itself
oxidised

Rusting is the oxidation of iron
by oxygen.

Exercises 2.1-2.2

2.1 Calculate the oxidation number of (a) S in SO_4^{2-} and (b) N in NO_3^- . [2]

2.2 Complete: “Oxidation is the _____ of electrons; the oxidation number goes _____.” [2]

Exercises 2.1·2.2

2.1 Calculate the oxidation number of (a) S in SO_4^{2-} and (b) N in NO_3^- . [2]

2.1

(a)

- S: $x + 4(-2) = -2 \Rightarrow x = +6$

(b)

- N: $x + 3(-2) = -1 \Rightarrow x = +5$

2.2 Complete: “Oxidation is the _____ of electrons; the oxidation number goes _____.” [2]

2.2

- loss; up

Exercises 3.1·3.2

3.1 What is the oxidation number of chromium in $\text{Cr}_2\text{O}_7^{2-}$?

[1]

☐

A +3

B +6

C +7

D +12

3.2 In the reaction $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$, the reducing agent is:

[1]

☐

A Zn

B Cu^{2+}

C Zn^{2+}

D Cu

Exercises 3.1·3.2

3.1 What is the oxidation number of chromium in $\text{Cr}_2\text{O}_7^{2-}$?

[1]

- ☐ **A** +3
B +6
C +7
D +12

3.1 B

- $2x + 7(-2) = -2 \Rightarrow 2x = 12 \Rightarrow x = +6$

3.2 In the reaction $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$, the reducing agent is:

[1]

- ☐ **A** Zn
B Cu^{2+}
C Zn^{2+}
D Cu

3.2 A

- Zn loses electrons (is oxidised), so it is the reducing agent

Exercise 3.3 (a)

3.3 When chlorine reacts with cold water it forms Cl^- and ClO^- .

(a) State the oxidation number of chlorine in Cl_2 , in Cl^- and in ClO^- .

[3]

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[3]

(a)

- $\text{Cl}_2 = 0$; $\text{Cl}^- = -1$; $\text{ClO}^- = +1$

Exercise 3.3 (b)

3.3 When chlorine reacts with cold water it forms Cl^- and ClO^- .

(b) Explain why this is a disproportionation reaction.

[2]

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(b) Explain why this is a disproportionation reaction.

[2]

(b)

- the same element (chlorine) is both oxidised (0
- +1) and reduced (0
- -1) in one reaction

Exercise 3.4

3.4 In acidic solution, MnO_4^- ($\text{Mn} = +7 \rightarrow +2$) oxidises Fe^{2+} ($\rightarrow \text{Fe}^{3+}$). Use oxidation-number changes to deduce the ratio of MnO_4^- to Fe^{2+} . [2]

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3.4

- Mn falls by 5 (+7
- +2); Fe rises by 1 (+2
- +3); so 5 Fe^{2+} are needed per MnO_4^- – ratio 1 : 5

Exercise 3.5 (a)

3.5 Chlorine reacts with **cold dilute** sodium hydroxide:



(a) Give the oxidation number of chlorine in Cl_2 , in NaCl and in NaClO .

[3]

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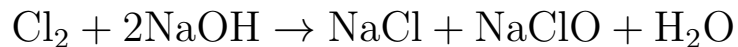
(a) Give the oxidation number of chlorine in Cl_2 , in NaCl and in NaClO . [3]

(a)

- Cl_2 : 0; NaCl : -1 ; NaClO : $+1$. In NaClO , Na is $+1$ and O is -2 , so Cl must be $+1$

Exercise 3.5 (b)

3.5 Chlorine reacts with **cold dilute** sodium hydroxide:



(b) Explain why this reaction is described as disproportionation.

[2]

Exercise 3.5 (b)

3.5 Chlorine reacts with **cold dilute** sodium hydroxide:



(b) Explain why this reaction is described as disproportionation. [2]

(b)

- The **same** element is oxidised and reduced in the one reaction: chlorine goes $0 \rightarrow +1$ in NaClO and $0 \rightarrow -1$ in NaCl

Exercise 3.5 (c)

3.5 Chlorine reacts with **cold dilute** sodium hydroxide:



(c) With **hot concentrated** sodium hydroxide the products are NaCl and NaClO₃. Write the balanced equation and show that this reaction is also disproportionation. [4]

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3.5 Chlorine reacts with **cold dilute** sodium hydroxide:



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(c)

- $3\text{Cl}_2 + 6\text{NaOH} \rightarrow 5\text{NaCl} + \text{NaClO}_3 + 3\text{H}_2\text{O}$. Five chlorine atoms fall $0 \rightarrow -1$ (gaining 5 electrons in total) while one rises $0 \rightarrow +5$ (losing 5); the electrons balance, and again one element is both oxidised and reduced

Exercise 3.5 (d)

3.5 Chlorine reacts with **cold dilute** sodium hydroxide:



(d) Chlorine also displaces iodine from aqueous potassium iodide. Write the ionic equation, and identify the oxidising agent, explaining your choice. [3]

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(d)

- $\text{Cl}_2 + 2\text{I}^- \rightarrow 2\text{Cl}^- + \text{I}_2$. Chlorine is the oxidising agent: it **gains** electrons ($0 \rightarrow -1$), i.e. it is itself reduced while oxidising the iodide