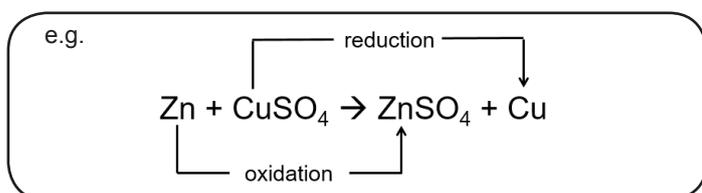


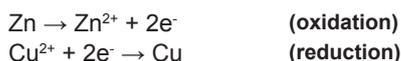
## Balancing Redox Equations in Acidic Solution

### Introduction

Redox is a portmanteau, i.e., a word blending the sounds and combining the meanings of two words; reduction and oxidation. Redox is defined as the loss of electrons by one chemical species (oxidation), and the gain of electrons by another chemical species (reduction).



This is illustrated using ionic half-equations:

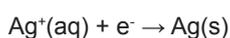


### Half-Equations

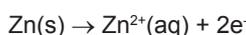
When completing equations for redox reactions, the oxidation and reduction processes are treated individually using separate ionic half-equations. These describe; (i) the oxidised chemical species, and, (ii) the reduced chemical species.

#### Example 1:

Silver ions (in aqueous solution) oxidise metallic zinc.  
The ionic half-equation for the reduction of silver ions:



The ionic half-equation for the oxidation of zinc:



Comparing the two half-equations reveals an inconsistency in the number of electrons involved. Before combining the half-equations, determine the correct factor to ensure the number of electrons gained by silver ions is equivalent to the number of electrons released by zinc atoms. In this example the factor is 2.



Combining the two half-equations provides a correctly balanced redox equation. When combining two half-equations, the number of electrons on both sides of the equation are omitted, i.e. they cancel out:



Giving the completed equation:



 Ionic equations are used to complete redox equations. Only chemical species undergoing a change in the reaction are included. Chemical species that remain unchanged in solution are called spectator ions.

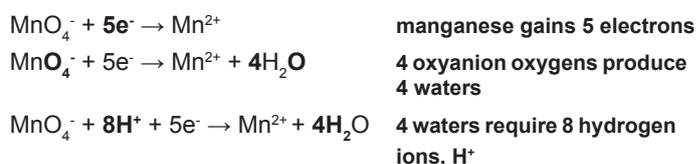
Half-equations involving oxyanions, i.e., anions consisting of a chemical element and oxygen, are more complicated. Typically, oxyanion oxygen atoms combine with hydrogen ions, H<sup>+</sup>, present in the reaction mixture, forming water, H<sub>2</sub>O. The hydrogen ions, H<sup>+</sup>, are usually provided by adding an excess of a strong acid to the reaction, for example, sulphuric acid, H<sub>2</sub>SO<sub>4</sub>, or nitric acid, HNO<sub>3</sub>.

### REDOX REACTIONS IN ACIDIC SOLUTIONS

Potassium manganate(VII), KMnO<sub>4</sub>, is a strong oxidising agent and a common reagent in redox reactions. In acidic conditions, manganese atoms in the +7 oxidation state gain five electrons and are reduced to the +2 oxidation state. The four oxyanion oxygens involved combine with eight hydrogen ions, H<sup>+</sup>, producing four water molecules.

#### Example 2:

The manganate(VII) ion, MnO<sub>4</sub><sup>-</sup>, is reduced to the manganese(II) ion, Mn<sup>2+</sup>.



Giving the completed half-equation:



### Balancing Redox Reactions In Acidic Solutions

When balancing redox equations, ascertain the reaction taking place, identifying the chemical species; (1) oxidised, and (2) reduced. Complete ionic half-equations, following the steps provided:

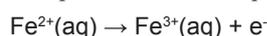
1. Confirm oxidation states/changes in oxidation states.
2. Ensure all non-oxygen/hydrogen atoms are balanced in the half-equation.
3. Determine the number of electrons; (1) lost through oxidation, and, (2) gained through reduction.
4. Accommodate any required oxygen atoms (present in oxyanions) by adding the required number of water molecules.
5. Accommodate any required hydrogen atoms by adding the required number of hydrogen ions.

#### Example 3a:

Potassium dichromate(VI), K<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub>, oxidises iron(II) ions to iron(III) ions under acidic conditions; dichromate(VI) ions are reduced to chromium(III).

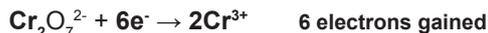
Iron(II) is oxidised to iron(III) – losing one electron.

Completed ionic half-equation for the oxidation of iron(II) ions:



## 273. Balancing Redox Equations in Acidic Solution

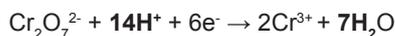
The dichromate(VI) ion,  $\text{Cr}_2\text{O}_7^{2-}$ , is reduced to the chromium(III) ion,  $2\text{Cr}^{3+}$ .



Ensure chromium atoms are balanced in the half-equation. There are **2** chromium atoms and each gains **3** electrons. Therefore, a total of **6 electrons** are required.

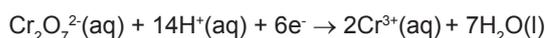


**7 oxyanion oxygens produce 7 waters**



**7 waters require 14 hydrogen ions,  $\text{H}^+$**

Completed ionic half-equation for the reduction of chromium:



The above example demonstrates the importance in ensuring non-oxygen/hydrogen atoms are correctly balanced. A common error occurs when this step is ignored and two half-equations are subsequently merged with incorrect stoichiometry.



*To confirm equations are correctly balanced, check that the net charge on both sides of the equation is equal. For example, in:*



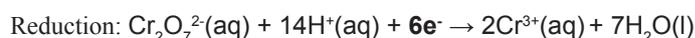
*The net charge on both sides of the equation is 6+.*

To correctly merge and complete the full redox equation:

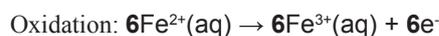
6. Determine the correct factor to balance the equations so the number of electrons is equal.
7. Merge (add) the two half-equations together.
8. Cancel the electrons (appearing on both sides of the merged equation).
9. If required, cancel any excess water molecules and/or hydrogen ions appearing on both sides of the merged equation (see **Example 4**).

### Example 3b:

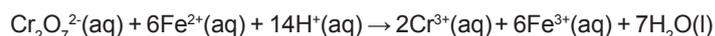
Compare the two redox half-equations completed above (see **Example 3a**), and factor the stoichiometric ratio to ensure the number of electrons gained in the reduction process matches the electrons released through oxidation.



Oxidation process requires the loss of 6 electrons, i.e., a factor of 6 is required:



Adding/merging both half-equations completes the fully balanced ionic equation:



As before, the electrons 'cancel out' and are not included in the final equation.

**Example 4** demonstrates how the number of hydrogen ions,  $\text{H}^+$ , and/or water molecules,  $\text{H}_2\text{O}$ , appearing on both sides of the combined equation are cancelled in the full equation.

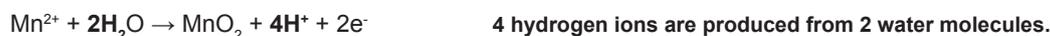
### Example 4:

Acidified potassium dichromate oxidises manganese(II) ions,  $\text{Mn}^{2+}$ , to manganese(IV),  $\text{MnO}_2$ ; dichromate(VI),  $\text{Cr}_2\text{O}_7^{2-}$ , is reduced to chromium(III),  $\text{Cr}^{3+}$ .

Refer back to **Example 3a** to complete the reduction half-equation:



Manganese(II) is oxidised to manganese(IV):



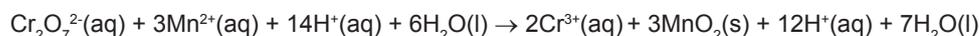
Completed ionic half-equation for the oxidation of manganese(II):



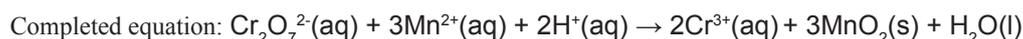
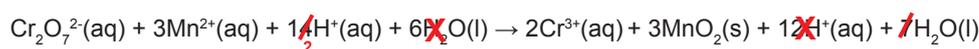
Factor the stoichiometry for both half-equations: 6 electrons gained by chromium atoms require 6 electrons released by manganese atoms, i.e., a factor of 3:



Merging both half-equations completes the fully balanced ionic equation:



The above equation contains hydrogen ions,  $\text{H}^+$ , and water molecules,  $\text{H}_2\text{O}$ , on both sides of the equation. The equation is simplified by cancelling excess hydrogen ions,  $\text{H}^+$ , and water molecules,  $\text{H}_2\text{O}$ :

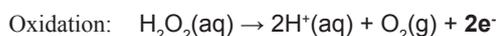


## 273. Balancing Redox Equations in Acidic Solution

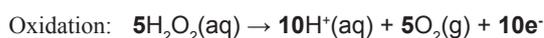
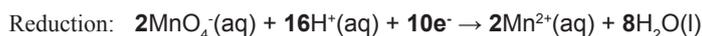
When determining the stoichiometric ratio/factor, apply simple whole-numbers, i.e., use the lowest common multiple to correctly balance half-equations. See **Example 5**.

### Example 5:

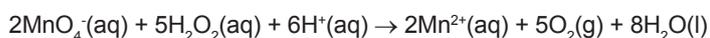
The oxidation of hydrogen peroxide,  $\text{H}_2\text{O}_2$ , by potassium manganate(VII),  $\text{KMnO}_4$ , gives the following redox half-equations:



The ratio observed is  $2\frac{1}{2} : 1$ . For simplicity, find the lowest common multiple and use whole-numbers, i.e., 10;  $(2 \times 5)$ ,  $(5 \times 2)$ :



Merging the half-equations, cancelling electrons and excess hydrogen ions,  $\text{H}^+$ , gives:

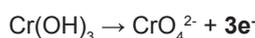


### Redox Reactions In Alkaline Solutions

Redox reactions occur in alkaline solutions. The principles for balancing redox equations in alkaline conditions are like those used completing in acidic conditions. The method is more complicated due to the required manipulation of equations using hydroxide ions,  $\text{OH}^-$ .

### Example 6:

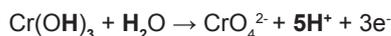
The oxidation of chromium(III) hydroxide,  $\text{Cr}(\text{OH})_3$ , to potassium chromate(VI),  $\text{K}_2\text{CrO}_4$ , is achieved using the oxidizing agent potassium chlorate(V),  $\text{KClO}_3$ . In the reaction, chlorate(V) is reduced to the chloride ion,  $\text{Cl}^-$ . This reaction is completed in alkaline conditions. Chromium(III) is oxidised to chromate(VI) – releasing 3 electrons.



**chromium loses 3 electrons.**

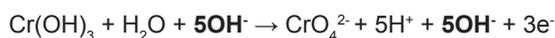


**balance oxygen atoms by adding  $\text{H}_2\text{O}$ .**

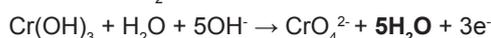
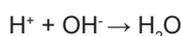


**balance hydrogen atoms by adding  $5\text{H}^+$ .**

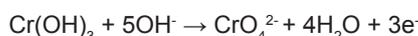
The reaction is completed in alkaline conditions. Any hydrogen ions,  $\text{H}^+$ , need to be countered by hydroxide ions,  $\text{OH}^-$ . It is necessary to add required hydroxide ions,  $\text{OH}^-$ , to both sides of the equation:



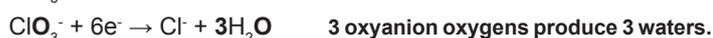
Hydrogen ions and hydroxide ions combine to form water:



Cancel water molecules,  $\text{H}_2\text{O}$ , present on both sides of the equation. This completes the oxidation half-equation:



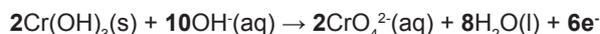
Chlorate(V) is reduced to chloride (-1) – gaining 6 electrons.



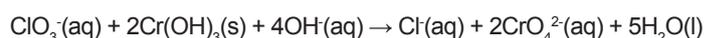
Add hydroxide ions,  $\text{OH}^-$ , to counter the hydrogen ions,  $\text{H}^+$ , on the left-side of the equation. Then, cancel excess water molecules appearing on both sides of the equation:



Balance the two half-equations. The factor in this example is 2.



Merge the two half-equations, cancelling electrons, hydroxide ions,  $\text{OH}^-$ , and water molecules,  $\text{H}_2\text{O}$ , appearing on both sides of the merged redox equation:



### Questions

- For each example provided, complete a balanced ionic half-equation:
  - copper is oxidised to copper(II) ions.
  - dioxovanadium(V) ions,  $\text{VO}_2^+$ , reduced to vanadium(II) ions,  $\text{V}^{2+}$
  - aqueous bromine is reduced to bromide ions.
  - bromine is oxidised to bromate(V).
- Potassium manganate(VII) is an oxidising agent. It oxidises vanadium(II),  $\text{V}^{2+}$ , to vanadate(V) ions,  $\text{VO}_3^-$ . This reaction is completed under acidic conditions.
  - Complete a balanced half-equation for the reduction of manganate(VII).
  - Complete a balanced half-equation for the oxidation of vanadium(II).
  - Give a balanced redox equation for this reaction.
  - Give the generic name used to describe the role of the potassium ions,  $\text{K}^+$ , in this reaction.
  - Sulphuric acid,  $\text{H}_2\text{SO}_4$ , is used to acidify the conditions for this reaction. Suggest why hydrochloric acid should NOT be used. Provide an ionic equation in support of this answer.
- Zinc metal reduces acidified dichromate(VI) ions,  $\text{Cr}_2\text{O}_7^{2-}$ , to form a solution containing zinc(II) ions,  $\text{Zn}^{2+}$ , and chromium(III) ions,  $\text{Cr}^{3+}$ .
  - Complete a balanced half-equation for the reduction process.
  - Complete a balanced half-equation for the oxidation process.
  - Give a balanced redox equation for this reaction.
- Aldehydes are oxidised to carboxylic acids using strong oxidising agents. For example, acidified potassium dichromate(VI) oxidises ethanal,  $\text{CH}_3\text{CHO}$ , to ethanoic acid,  $\text{CH}_3\text{COOH}$ . Typically, dichromate(VI),  $\text{Cr}_2\text{O}_7^{2-}$ , is reduced to chromium(III),  $\text{Cr}^{3+}$ .
  - Complete the half-equation for the reduction process.
  - Complete the incomplete half-equation describing the oxidation of ethanal:  $\text{C}_2\text{H}_4\text{O} \rightarrow \text{C}_2\text{H}_4\text{O}_2$
  - Give a balanced redox equation for this reaction.
- Aluminium-oxygen fuel cells are being developed that may potentially power motor vehicles. This redox reaction occurs under alkaline conditions.
  - Complete the incomplete half-equation describing the oxidation of aluminium:  $\text{Al}(\text{s}) \rightarrow \text{Al}(\text{OH})_4^-(\text{aq})$
  - Complete the incomplete half-equation for the reduction of oxygen:  $\text{O}_2(\text{g}) \rightarrow 4\text{OH}^-(\text{aq})$
  - Give the overall equation for the aluminium-oxygen cell.

273. Balancing Redox Equations in Acidic Solution

**Answers**

1. (a)  $\text{Cu(s)} \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$   
 (b)  $\text{VO}_2^+(\text{aq}) + 4\text{H}^+(\text{aq}) + 3\text{e}^- \rightarrow \text{V}^{2+}(\text{aq}) + 2\text{H}_2\text{O(l)}$   
 (c)  $\text{Br}_2(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Br}^-(\text{aq})$   
 (d)  $\text{Br}_2 + 6\text{H}_2\text{O(l)} \rightarrow 2\text{BrO}_3^- + 12\text{H}^+(\text{aq}) + 10\text{e}^-$
  
2. (a)  $\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O(l)}$   
 (b)  $\text{V}^{2+}(\text{aq}) + 3\text{H}_2\text{O} \rightarrow \text{VO}_3^-(\text{aq}) + 6\text{H}^+(\text{aq}) + 3\text{e}^-$   
 (c)  $3\text{MnO}_4^-(\text{aq}) + 5\text{V}^{2+}(\text{aq}) + 3\text{H}_2\text{O(l)} \rightarrow 3\text{Mn}^{2+}(\text{aq}) + 5\text{VO}_3^-(\text{aq}) + 6\text{H}^+(\text{aq})$   
 (d) Spectator ion.  
 (e) Toxic chlorine gas is liberated.  $2\text{Cl}^-(\text{aq}) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$ , or  
 $2\text{MnO}_4^-(\text{aq}) + 10\text{Cl}^-(\text{aq}) + 16\text{H}^+(\text{aq}) \rightarrow 2\text{Mn}^{2+}(\text{aq}) + 5\text{Cl}_2(\text{g}) + 8\text{H}_2\text{O(l)}$
  
3. (a)  $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 8\text{e}^- \rightarrow 2\text{Cr}^{2+}(\text{aq}) + 7\text{H}_2\text{O(l)}$   
 (b)  $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$   
 (c)  $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 4\text{Zn(s)} + 14\text{H}^+(\text{aq}) \rightarrow 2\text{Cr}^{2+}(\text{aq}) + 4\text{Zn}^{2+}(\text{aq}) + 7\text{H}_2\text{O(l)}$
  
4. (a)  $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O(l)}$   
 (b)  $\text{C}_2\text{H}_4\text{O(l)} + \text{H}_2\text{O(l)} \rightarrow \text{C}_2\text{H}_4\text{O}_2(\text{l}) + 2\text{H}^+(\text{aq}) + 2\text{e}^-$   
 (c)  $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 3\text{C}_2\text{H}_4\text{O(l)} + 8\text{H}^+(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 3\text{C}_2\text{H}_4\text{O}_2(\text{l}) + 4\text{H}_2\text{O(l)}$
  
5. (a)  $\text{Al(s)} + 4\text{OH}^-(\text{aq}) \rightarrow \text{Al(OH)}_4^-(\text{aq}) + 3\text{e}^-$   
 (b)  $\text{O}_2(\text{g}) + 2\text{H}_2\text{O(l)} + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$   
 (c)  $3\text{O}_2(\text{g}) + 4\text{Al(s)} + 4\text{OH}^-(\text{aq}) + 6\text{H}_2\text{O(l)} \rightarrow 4\text{Al(OH)}_4^-(\text{aq})$