12. REDOX EQUILIBRIA

The electrochemical series (reference table)
12.1. Redox reactions
12.2. Standard electrode potentials
12.3. Calculations involving electrochemical cells
12.4. Using $E^\circ$ values to predict reactions
12.5. Applications of electrochemical cells

Redox equilibria answers
The electrochemical series

Good **reducing agents** (good at losing electrons)

<table>
<thead>
<tr>
<th>Reduction half equation</th>
<th>( E^\circ / V )</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{Li}^+(aq) + 1 \text{e}^- \rightleftharpoons \text{Li(s)} )</td>
<td>-3.03</td>
</tr>
<tr>
<td>( \text{Ba}^{2+}(aq) + 2 \text{e}^- \rightleftharpoons \text{Ba(s)} )</td>
<td>-2.90</td>
</tr>
<tr>
<td>( \text{Ca}^{2+}(aq) + 2 \text{e}^- \rightleftharpoons \text{Ca(s)} )</td>
<td>-2.87</td>
</tr>
<tr>
<td>( \text{Al}^{3+}(aq) + 2 \text{e}^- \rightleftharpoons \text{Al(s)} )</td>
<td>-1.66</td>
</tr>
<tr>
<td>( \text{Zn}^{2+}(aq) + 2 \text{e}^- \rightleftharpoons \text{Zn(s)} )</td>
<td>-0.76</td>
</tr>
<tr>
<td>( \text{Fe}^{2+}(aq) + 2 \text{e}^- \rightleftharpoons \text{Fe(s)} )</td>
<td>-0.44</td>
</tr>
<tr>
<td>( \text{Cr}^{3+}(aq) + 1 \text{e}^- \rightleftharpoons \text{Cr}^{2+}(aq) )</td>
<td>-0.41</td>
</tr>
<tr>
<td>( \text{V}^{3+}(aq) + 1 \text{e}^- \rightleftharpoons \text{V}^{2+}(aq) )</td>
<td>-0.26</td>
</tr>
<tr>
<td>( \text{Sn}^{2+}(aq) + 2 \text{e}^- \rightleftharpoons \text{Sn(s)} )</td>
<td>-0.14</td>
</tr>
<tr>
<td>( 2 \text{H}^+(aq) + 2 \text{e}^- \rightleftharpoons \text{H}_2(aq) )</td>
<td>0.00</td>
</tr>
<tr>
<td>( \text{Cu}^{2+}(aq) + 1 \text{e}^- \rightleftharpoons \text{Cu}^{+}(aq) )</td>
<td>+0.15</td>
</tr>
<tr>
<td>( \text{Cu}^{2+}(aq) + 2 \text{e}^- \rightleftharpoons \text{Cu(s)} )</td>
<td>+0.34</td>
</tr>
<tr>
<td>( \text{VO}^{2+}(aq) + 2 \text{H}^+(aq) + 1 \text{e}^- \rightleftharpoons \text{V}^{3+}(aq) + \text{H}_2\text{O(l)} )</td>
<td>+0.34</td>
</tr>
<tr>
<td>( \text{Cu}^{+}(aq) + 1 \text{e}^- \rightleftharpoons \text{Cu(s)} )</td>
<td>+0.52</td>
</tr>
<tr>
<td>( \text{I}_2(s) + 2 \text{e}^- \rightleftharpoons 2 \text{I}^-(aq) )</td>
<td>+0.54</td>
</tr>
<tr>
<td>( \text{MnO}_4^-(aq) + 1 \text{e}^- \rightleftharpoons \text{MnO}_4^{2-}(aq) )</td>
<td>+0.56</td>
</tr>
<tr>
<td>( \text{MnO}_4^{2-}(aq) + 2 \text{H}_2\text{O(l)} + 2 \text{e}^- \rightleftharpoons \text{MnO}_2(s) + 4 \text{OH}^-(aq) )</td>
<td>+0.59</td>
</tr>
<tr>
<td>( \text{O}_2(g) + 2 \text{H}^+(aq) + 2 \text{e}^- \rightleftharpoons \text{H}_2\text{O}_2(aq) )</td>
<td>+0.68</td>
</tr>
<tr>
<td>( \text{Fe}^{3+}(aq) + 1 \text{e}^- \rightleftharpoons \text{Fe}^{2+}(aq) )</td>
<td>+0.77</td>
</tr>
<tr>
<td>( \text{Ag}^+(aq) + 1 \text{e}^- \rightleftharpoons \text{Ag(s)} )</td>
<td>+0.80</td>
</tr>
<tr>
<td>( \text{VO}_2^+(aq) + 2 \text{H}^+(aq) + 1 \text{e}^- \rightleftharpoons \text{VO}^{2+}(aq) + \text{H}_2\text{O(l)} )</td>
<td>+1.00</td>
</tr>
<tr>
<td>( \text{Br}_2(g) + 2 \text{e}^- \rightleftharpoons 2 \text{Br}^-(aq) )</td>
<td>+1.07</td>
</tr>
<tr>
<td>( \frac{1}{2} \text{O}_2(g) + 2 \text{H}^+ + 2 \text{e}^- \rightleftharpoons \text{H}_2\text{O(l)} )</td>
<td>+1.23</td>
</tr>
<tr>
<td>( \text{Cr}_2\text{O}_7^{2-}(aq) + 14 \text{H}^+(aq) + 6 \text{e}^- \rightleftharpoons 2 \text{Cr}^{3+}(aq) + 7 \text{H}_2\text{O(l)} )</td>
<td>+1.33</td>
</tr>
<tr>
<td>( \text{Cl}_2(g) + 2 \text{e}^- \rightleftharpoons 2 \text{Cl}^-(aq) )</td>
<td>+1.36</td>
</tr>
<tr>
<td>( \text{MnO}_2(aq) + 8 \text{H}^+(aq) + 5 \text{e}^- \rightleftharpoons \text{Mn}^{2+}(aq) + 4 \text{H}_2\text{O(l)} )</td>
<td>+1.51</td>
</tr>
<tr>
<td>( \text{HClO}(g) + \text{H}^+(aq) + 1 \text{e}^- \rightleftharpoons \frac{1}{2} \text{Cl}_2(g) + \text{H}_2\text{O(l)} )</td>
<td>+1.59</td>
</tr>
<tr>
<td>( \text{H}_2\text{O}_2(aq) + 2 \text{H}^+(aq) + 2 \text{e}^- \rightleftharpoons 2 \text{H}_2\text{O(l)} )</td>
<td>+1.77</td>
</tr>
<tr>
<td>( \text{F}_2(g) + 2 \text{e}^- \rightleftharpoons 2 \text{F}^-(aq) )</td>
<td>+2.87</td>
</tr>
</tbody>
</table>

Good **oxidising agents** (good at gaining electrons)
12.1. Redox reactions

Faisal has written the following notes on redox reactions in preparation for his AS exams. However there are a few mistakes, many of which are commonly seen in exam answers. Help Faisal learn from his mistakes by correcting the errors so that he has an accurate set of notes to revise from. (10 marks)

**Redox reactions**

Oxidation is the loss of electrons; the oxidation state decreases

Reduction is the gain of electrons; the oxidation state increases

An oxidising agent oxidises something and is itself reduced

A reducing agent reduces something and is itself oxidised

<table>
<thead>
<tr>
<th>Oxidation</th>
<th>Reduction</th>
</tr>
</thead>
<tbody>
<tr>
<td>+3</td>
<td>+1</td>
</tr>
<tr>
<td>Fe₂O₃</td>
<td>Fe</td>
</tr>
<tr>
<td>+1</td>
<td>+2</td>
</tr>
<tr>
<td>CO</td>
<td>CO₂</td>
</tr>
</tbody>
</table>

Fe₂O₃ + 3 CO → 2 Fe + 3 CO₂

**Rules for balancing redox reactions**

1. Use oxidation states to identify the species undergoing oxidation or reduction.
   
   Fe₂O₃ → Fe (reduction)  CO → CO₂ (oxidation)

2. Balance the atoms undergoing oxidation or reduction.
   
   Fe₂O₃ → 2 Fe  CO → CO₂

3. Balance the oxygen atoms by adding H₂O.
   
   Fe₂O₃ → 2 Fe + 3 H₂O  H₂O + CO → CO₂

4. Balance the hydrogen atoms by adding H⁺.
   
   6 H⁺ + Fe₂O₃ → 2 Fe + 3 H₂O  H₂O + CO → CO₂ + 2 H⁺

5. Balance the electrons.
   
   6 e⁻ + 6 H⁺ + Fe₂O₃ → 2 Fe + 3 H₂O  2 e⁻ + H₂O + CO → CO₂ + 2 H⁺

6. Multiply either half equation to make sure that the no. of electrons in each half equation match.
   
   6 e⁻ + 6 H⁺ + Fe₂O₃ → 2 Fe + 3 H₂O  6 e⁻ + 3 H₂O + 3 CO → CO₂ + 6 H⁺

7. Combine to get the full redox equation.
   
   6 e⁻ + 6 H⁺ + Fe₂O₃ + 3 H₂O + 3 CO → 2 Fe + 3 H₂O + 3 CO₂ + 6 H⁺ + 6 e⁻

8. Cancel out anything that appears on both sides of the equation.
   
   6 e⁻ + 6 H⁺ + Fe₂O₃ + 3 H₂O + 3 CO → 2 Fe + 3 H₂O + 3 CO₂ + 6 H⁺ + 6 e⁻

Fe₂O₃ + 3 CO → 2 Fe + 3 CO₂

Redox equilibria 12.1.
12.2. Standard electrode potentials

We can measure how readily something gives away electrons by measuring its standard electrode potential, $E^\circ$.

1. Standard electrode potentials are measured by connecting a half cell containing the equilibrium, the potential of which is to be measured to a standard hydrogen electrode at 298 K.
   (a) Label the diagram below showing the standard hydrogen electrode. (3 marks)
   (b) Complete the diagram to show the complete cell you would use if you wished to measure $E^\circ$ for a zinc electrode. (4 marks)

![Diagram]

2. Cells can be represented in shorthand form using a series of standard conventions.
   (a) Match up the symbol to its meaning when used to represent an electrochemical cell;
       | Shows a salt bridge
       || Indicates a phase boundary (1 mark)
   (b) For each half cell, the species in the highest oxidation state in the redox equilibrium is written next to the salt bridge.
       Use this convention to complete the shorthand representation of the cells produced when half cells containing each of the equilibria below are connected to a standard hydrogen electrode.
       (i) $\text{Fe}^{2+}(aq) + 2 \ e^- \rightleftharpoons \text{Fe}(s)$; $\text{Pt} | H_2(g) | H^+(aq)$ .................................................................(1 mark)
       (ii) $\text{MnO}_4^-(aq) + 1 \ e^- \rightleftharpoons \text{MnO}_4^{2-}(aq)$; $\text{Pt} | H_2(g) | H^+(aq)$ .................................................................(1 mark)
1. For each of the electrochemical cells described below;
   (i) Calculate the emf of the cell as written,
   (ii) Identify the reaction occurring at the positive and negative electrodes,
   (iii) Write an equation for the overall cell reaction which occurs when the
         electrodes are connected

   Assume standard conditions. Use the table of standard electrode potentials provided at the start of this
   chapter for reference.

   (a) \( \text{Al(s)} | \text{Al}^{3+}(aq) \parallel \text{Zn}^{2+}(aq) | \text{Zn(s)} \)

   \[ E^{\circ}_{\text{cell}} = \ \text{Calculate the emf of the cell as written,} \]

   \[ \text{Positive electrode half equation} \]

   \[ \text{Negative electrode half equation} \]

   \[ \text{Overall cell reaction} \]

   (4 marks)

   (b) \( \text{Cl}^{-}(aq) | \text{Cl}_2(g) \parallel \text{Fe}^{2+}(aq) | \text{Fe(s)} \)

   \[ E^{\circ}_{\text{cell}} = \ \text{Calculate the emf of the cell as written,} \]

   \[ \text{Positive electrode half equation} \]

   \[ \text{Negative electrode half equation} \]

   \[ \text{Overall cell reaction} \]

   (4 marks)

2. The electrochemical cell shown below is set up;
   \[ \text{Pt} | \text{Mn}^{2+}(aq), \text{MnO}_2(aq) \parallel \text{S}_2\text{O}_8^{2-}(aq), \text{SO}_4^{2-}(aq) | \text{Pt}, E^{\circ}_{\text{cell}} = +0.50 \text{V} \]

   (a) Calculate the standard electrode potential for the following half-reaction;

   \[ \text{S}_2\text{O}_8^{2-}(aq) + 2 e^- \rightleftharpoons 2 \text{SO}_4^{2-}(aq) \]

   (1 mark)

   (b) For the standard electrode potentials, all ion concentrations must be 1 mol dm\(^{-3}\).

   Deduce what effect an increase in the concentration of \( \text{S}_2\text{O}_8^{2-}(aq) \) ions to higher than 1 mol dm\(^{-3}\)
   would have on \( E^{\circ}_{\text{cell}} \).

   (1 mark)
12.4. Using $E_{\text{m sat}}$ values to predict reactions

We can use standard electrode potentials to predict if reactions will happen. Remember the more negative the electrode potential the better the species on the right of the reduction half equation is as a reducing agent.

1. (a) Predict if the following reactions are feasible or not;
   (i) \( \text{Cu}^{2+}(\text{aq}) + 2 \text{Br}^- (\text{aq}) \rightarrow \text{Cu}(s) + \text{Br}_2(g) \) .................................................. (1 mark)
   (ii) \( \text{Fe}(s) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{Cu}(s) \) ..................................................(1 mark)

(b) Choose a halogen that could be used to oxidise \( \text{Fe}^{2+}(\text{aq}) \) ions to \( \text{Fe}^{3+}(\text{aq}) \) ions. Explain your choice.

(c) When carrying out oxidation reactions, a solution of \( \text{MnO}_4^- (\text{aq}) \) cannot be acidified by HCl whereas a solution of \( \text{Cr}_2\text{O}_7^{2-} \) in theory can.

   Explain why, using your understanding of standard electrode potentials.

2. For each of the following combination of reagents, choose from the options below the final colour of the solution. Write an equation to represent the overall reaction occurring.

   - green
   - blue
   - yellow
   - orange
   - lilac

(a) Addition of tin powder to a solution of \( \text{VO}_2^+ \) ions in acidic conditions.
   \( [\text{V}^{2+}(\text{aq}) \text{ is lilac; } \text{V}^{3+}(\text{aq}) \text{ is green; } \text{VO}^{2+}(\text{aq}) \text{ is blue and } \text{VO}_2^+(aq) \text{ is yellow}] \)

(b) Addition of silver powder to a solution of dichromate \( \text{(Cr}_2\text{O}_7^{2-} \) ions in acidic conditions.
   \( [\text{Cr}_2\text{O}_7^{2-}(aq) \text{ is orange; } \text{Cr}^{3+}(aq) \text{ is green; } \text{Cr}^{2+}(aq) \text{ is blue}] \)

Redox equilibria 12.4.
12.5. Applications of electrochemical cells

Technically a **battery is two or more simple cells connected together**. However in everyday speech we rarely make the distinction.

There are a wide variety of batteries available today. The table below shows some details of the chemistry involved in just a few.

<table>
<thead>
<tr>
<th>Battery type</th>
<th>Half equations with electrode potentials</th>
<th>( E_{\text{cell}} / \text{V} )</th>
</tr>
</thead>
</table>
| Zinc/carbon    | \( \text{Zn}^{2+}(aq) + 2 \text{e}^- \rightleftharpoons \text{Zn}(s) \)  
2 \( \text{NH}_3(aq) + 2 \text{e}^- \rightleftharpoons 2 \text{NH}_3(g) + \text{H}_2(g) \) | \( E \approx -0.8 \text{ V} \)  
\( E \approx +0.7 \text{ V} \) |
| Nickel/cadmium | \( \text{Cd(OH)}_2(s) + 2 \text{e}^- \rightleftharpoons \text{Cd}(s) + 2 \text{OH}^- (aq) \)  
\( \text{NiO(OH)}(s) + \text{H}_2\text{O(l)} + 1 \text{e}^- \rightleftharpoons \text{Ni(OH)}_2(s) + \text{OH}^- (aq) \) | \( E \approx -0.8 \text{ V} \)  
\( E \approx +0.5 \text{ V} \) |
| Lead-acid      | \( \text{PbSO}_4(s) + 2 \text{e}^- \rightleftharpoons \text{Pb}(s) + \text{SO}_4^{2-}(aq) \)  
\( \text{PbO}_2(s) + 4 \text{H}^+(aq) \rightleftharpoons \text{PbSO}_4(s) + 2 \text{H}_2\text{O(l)} \)  
+ \( \text{SO}_4^{2-}(aq) + 2 \text{e}^- \) | \( E \approx -0.35 \text{ V} \)  
\( E \approx +1.70 \text{ V} \) |
| Fuel cell      | \( 2 \text{H}^+(aq) + 2 \text{e}^- \rightleftharpoons \text{H}_2(g) \)  
\( 4 \text{H}^+(aq) + \text{O}_2(g) + 4 \text{e}^- \rightleftharpoons 2 \text{H}_2\text{O(l)} \) | \( E \approx 0.0 \text{ V} \)  
\( E \approx +1.2 \text{ V} \) |

1. Complete the table above by calculating the e.m.f. for each of the different cell types. (4 marks)

2. Consider the nickel/cadmium cell in more detail.
   (a) Identify the element which undergoes a change in oxidation state at the positive electrode and state the oxidation state change.

   ...................................................................................................................................................(2 marks)

   (b) Write the conventional representation of the cell.

   ...................................................................................................................................................(2 marks)

   (c) The nickel-cadmium cell is rechargeable. Write an equation for the overall reaction that occurs when the battery is being **recharged**.

   ..................................................................................................................................................

   ...................................................................................................................................................(1 mark)

   (d) Nickel/cadmium cells must be carefully disposed of. Suggest one reason why.

   ...................................................................................................................................................(1 mark)
# Redox equilibria answers

## 12.1. Redox reactions

Corrections shown in red and circled. (10 marks; 1 mark for each correct correction made)

<table>
<thead>
<tr>
<th>Redox reactions</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Oxidation</strong> is the loss of electrons; the oxidation state <strong>increases</strong></td>
</tr>
<tr>
<td><strong>Reduction</strong> is the gain of electrons; the oxidation state <strong>decreases</strong></td>
</tr>
<tr>
<td>An oxidising agent oxidises something and is itself reduced</td>
</tr>
<tr>
<td>A reducing agent reduces something and is itself oxidised</td>
</tr>
</tbody>
</table>

**Examples:**

- Fe\(_2\)O\(_3\) \(\rightarrow\) Fe \(\text{reduction}\)
- CO \(\rightarrow\) CO\(_2\) \(\text{oxidation}\)

## Rules for balancing redox reactions

1. Use oxidation states to identify the species undergoing oxidation or reduction.
   - Fe\(_2\)O\(_3\) \(\rightarrow\) Fe \(\text{reduction}\)
   - CO \(\rightarrow\) CO\(_2\) \(\text{oxidation}\)

2. Balance the atoms in each undergoing oxidation or reduction.
   - Fe\(_2\)O\(_3\) \(\rightarrow\) 2 Fe
   - CO \(\rightarrow\) CO\(_2\)

3. Balance the oxygen atoms by adding H\(_2\)O.
   - Fe\(_2\)O\(_3\) \(\rightarrow\) 2 Fe + 3 H\(_2\)O
   - H\(_2\)O + CO \(\rightarrow\) CO\(_2\)

4. Balance the hydrogen atoms by adding H\(^+\).
   - 6 H\(^+\) + Fe\(_2\)O\(_3\) \(\rightarrow\) 2 Fe + 3 H\(_2\)O
   - H\(_2\)O + CO \(\rightarrow\) CO\(_2\) + 2 H\(^+\)

5. Balance the electrons.
   - 6 e\(^-\) + 6 H\(^+\) + Fe\(_2\)O\(_3\) \(\rightarrow\) 2 Fe + 3 H\(_2\)O
   - H\(_2\)O + CO \(\rightarrow\) CO\(_2\) + 2 H\(^+\) + 2 e\(^-\)

6. Multiply either half equation to make sure that the no. of electrons in each half equation match.
   - 6 e\(^-\) + 6 H\(^+\) + Fe\(_2\)O\(_3\) \(\rightarrow\) 2 Fe + 3 H\(_2\)O
   - 3 H\(_2\)O + 3 CO \(\rightarrow\) 3 CO\(_2\) + 6 H\(^+\) + 6 e\(^-\)

7. Combine to get the full redox equation.
   - 6 e\(^-\) + 6 H\(^+\) + Fe\(_2\)O\(_3\) + 3 H\(_2\)O + 3 CO \(\rightarrow\) 2 Fe + 3 H\(_2\)O + 3 CO\(_2\) + 6 H\(^+\) + 6 e\(^-\)

8. Cancel out anything that appears on both sides of the equation.
   - 6 e\(^-\) + 6 H\(^+\) + Fe\(_2\)O\(_3\) + 3 H\(_2\)O + 3 CO \(\rightarrow\) 2 Fe + 3 H\(_2\)O + 3 CO\(_2\) + 6 H\(^+\) + 6 e\(^-\)
   - Fe\(_2\)O\(_3\) + 3 CO \(\rightarrow\) 2 Fe + 3 CO\(_2\)
12. Redox equilibria answers

12.2. Standard electrode potentials

1. (a) Shows a salt bridge ɳ Indicates a phase boundary (1 mark)

(b) (i) Pt | H₂(g) | H⁺(aq) Fe²⁺(aq) | Fe(s) (1 mark)

(ii) Pt | H₂(g) | H⁺(aq) MnO₄⁻(aq), MnO₄²⁻(aq) | Pt (1 mark)

2. (a) +0.50 V = E°<sub>RHS</sub> − (+1.51 V), ∴ E°<sub>RHS</sub> = +2.01 V (1 mark)

12.3. Calculations involving electrochemical cells

1. (a) E°<sub>cell</sub> = −0.76 − (−1.66) = +0.90 V (1 mark)

   Positive electrode: Zn²⁺(aq) + 2 e⁻ → Zn(s) × 3 (1 mark)

   Negative electrode: Al(s) → Al³⁺(aq) + 3 e⁻ × 2 (1 mark)

   Overall cell reaction: 3 Zn²⁺(aq) + 2 Al(s) → 3 Zn(s) + 2 Al³⁺(aq) (1 mark)

(b) E°<sub>cell</sub> = −0.44 − (+1.36) = −1.80 V (1 mark)

   Positive electrode: Cl₂(g) + 2 e⁻ → 2 Cl⁻(aq) (1 mark)

   Negative electrode: Fe(s) → Fe²⁺(aq) + 2 e⁻ (1 mark)

   Overall cell reaction: Cl₂(g) + Fe(s) → 2 Cl⁻(aq) + Fe²⁺(aq) (1 mark)

2. (a) +0.50 V = E°<sub>RHS</sub> − (+1.51 V), ∴ E°<sub>RHS</sub> = +2.01 V (1 mark)
12. Redox equilibria answers

(b) According to Le Chatelier’s principle, an increase in the concentration of $S_2O_8^{2-}$(aq) ions causes the equilibrium to shift to the right (using up electrons) and therefore $E^\circ_{\text{RHS}}$ will become more positive. Since $E^\circ_{\text{cell}} = E^\circ_{\text{RHS}} - E^\circ_{\text{LHS}}$, the more positive $E^\circ_{\text{RHS}}$ the more positive $E^\circ_{\text{cell}}$. Therefore $E^\circ_{\text{cell}}$ will increase / become more positive. (1 mark)

12.4. Using $E_{\text{red}}$ values to predict reactions

1. (a) (i) $\text{Cu}^{2+}$(aq) + 2 e$^-$ ⇌ Cu(s) $E^\circ = +0.34$ V, so Cu(s) → Cu$^{2+}$(aq) + 2 e$^-$

Br$_2$(g) + 2 e$^-$ ⇌ 2 Br$^-$(aq) $E^\circ = +1.07$ V, so Br$_2$(g) + 2 e$^-$ → 2 Br$^-$(aq) Hence the reaction, Cu$^{2+}$(aq) + 2 Br$^-$(aq) → Cu(s) + Br$_2$(g) is not feasible (1 mark)

(ii) $\text{Fe}^{2+}$(aq) + 2 e$^-$ ⇌ Fe(s) $E^\circ = -0.44$ V, so Fe(s) → Fe$^{2+}$(aq) + 2 e$^-$

Cu$^{2+}$(aq) + 2 e$^-$ ⇌ Cu(s) $E^\circ = +0.34$ V, so Cu$^{2+}$(aq) + 2 e$^-$ ⇌ Cu(s) Hence the reaction, Fe(s) + Cu$^{2+}$(aq) → Fe$^{2+}$(aq) + Cu(s) is feasible (1 mark)

(b) $\text{Fe}^{3+}$(aq) + 1 e$^-$ ⇌ Fe$^{2+}$(aq) $E^\circ = +0.77$ V

F$_2$(g) + 2 e$^-$ ⇌ 2 F$^-$(aq) $E^\circ = +2.87$ V

Cl$_2$(g) + 2 e$^-$ ⇌ 2 Cl$^-$(aq) $E^\circ = +1.36$ V

Br$_2$(g) + 2 e$^-$ ⇌ 2 Br$^-$(aq) $E^\circ = +1.07$ V

I$_2$(g) + 2 e$^-$ ⇌ 2 I$^-$(aq) $E^\circ = +0.54$ V

For the halogen to oxidise the Fe$^{3+}$ ions, +0.77 V must be the more negative reduction potential. Therefore any of fluorine, chlorine or bromine would be a suitable oxidising agent. (2 marks)

(c) Cr$_2$O$_7^{2-}$(aq) + 14 H$^+$ (aq) + 6 e$^-$ ⇌ 2 Cr$^{3+}$(aq) + 7 H$_2$O(l) $E^\circ = +1.33$ V

MnO$_2$(aq) + 8 H$^+$ (aq) + 5 e$^-$ ⇌ Mn$^{2+}$(aq) + 4 H$_2$O(l) $E^\circ = +1.51$ V

Cl$_2$(g) + 2 e$^-$ ⇌ 2 Cl$^-$(aq) $E^\circ = +1.36$ V

Comparing initially at the reduction potentials for MnO$_2$/Mn$^{2+}$ and Cl$_2$/Cl$^-$, the latter has the more negative value and hence the chloride ions in the hydrochloric acid will be oxidised by the MnO$_2$ to produce chlorine gas which is toxic. Hence a solution of MnO$_2$ cannot be acidified by HCl. (1 mark)

Comparing now at the reduction potentials for Cr$_2$O$_7^{2-}$/Cr$^{3+}$ and Cl$_2$/Cl$^-$, the formed now has the more positive value and hence the chloride ions cannot be oxidised by Cr$_2$O$_7^{2-}$ (this of course assumes standard conditions). (1 mark)

2. (a) Sn$^{2+}$(aq) + 2 e$^-$ ⇌ Sn(s) $E^\circ = -0.14$ V

V$^{2+}$(aq) + 1 e$^-$ ⇌ V$^{3+}$(aq) $E^\circ = -0.26$ V

Redox equilibria answers
12. Redox equilibria answers

\[ \text{VO}^{2+}(aq) + 2 \text{H}^+(aq) + 1 \text{e}^- \rightleftharpoons \text{V}^{3+}(aq) + \text{H}_2\text{O}(l) \quad \text{E}^\circ = +0.34 \text{ V} \]

\[ \text{VO}_2^+(aq) + 2 \text{H}^+(aq) + 1 \text{e}^- \rightleftharpoons \text{VO}^{2+}(aq) + \text{H}_2\text{O}(l) \quad \text{E}^\circ = +1.00 \text{ V} \]

Between \( \text{VO}_2^+/\text{VO}^{2+} \) and \( \text{Sn}^{2+}/\text{Sn} \) the latter has the more negative electrode potential and so \( \text{Sn} \) will act as a reducing agent, reducing \( \text{VO}_2^+ \) to \( \text{VO}^{2+} \):

\( \text{Sn}(s) \rightarrow \text{Sn}^{2+}(aq) + 2 \text{e}^- \) and \( \text{VO}_2^+(aq) + 2 \text{H}^+(aq) + 1 \text{e}^- \rightarrow \text{VO}^{2+}(aq) + \text{H}_2\text{O}(l) \)

Between \( \text{VO}^{2+/\text{V}^{3+}} \) and \( \text{Sn}^{2+}/\text{Sn} \) again the latter has the more negative electrode potential and so \( \text{Sn} \) will act as a reducing agent, reducing \( \text{VO}^{2+} \) to \( \text{V}^{3+} \):

\( \text{Sn}(s) \rightarrow \text{Sn}^{2+}(aq) + 2 \text{e}^- \) and \( \text{VO}^{2+}(aq) + 2 \text{H}^+(aq) + 1 \text{e}^- \rightarrow \text{V}^{3+}(aq) + \text{H}_2\text{O}(l) \)

Finally between \( \text{V}^{3+/\text{V}^{2+}} \) and \( \text{Sn}^{2+}/\text{Sn} \) the former has the more negative electrode potential and so no further reaction will occur.

The colour of the final solution will therefore be **green** owing to the presence of \( \text{V}^{3+}(aq) \) ions.

The overall equation for the reaction is:

\[ 2 \text{Sn}(s) + \text{VO}_2^+(aq) + 4 \text{H}^+ \rightarrow 2 \text{Sn}^{2+}(aq) + \text{V}^{3+}(aq) + 2 \text{H}_2\text{O}(l) \]

*(1 mark for the identification of the correct final colour, 1 mark for the overall equation)*

(b) \( \text{Ag}^+(aq) + 1 \text{e}^- \rightleftharpoons \text{Ag}(s) \quad \text{E}^\circ = +0.80 \text{ V} \)

\[ \text{Cr}_2\text{O}_7^{2-}(aq) + 14 \text{H}^+(aq) + 6 \text{e}^- \rightleftharpoons 2 \text{Cr}^{3+}(aq) + 7 \text{H}_2\text{O}(l) \quad \text{E}^\circ = +1.33 \text{ V} \]

\[ \text{Cr}^{3+}(aq) + 1 \text{e}^- \rightleftharpoons \text{Cr}^{2+}(aq) \quad \text{E}^\circ = -0.41 \text{ V} \]

Between \( \text{Ag}^+/\text{Ag} \) and \( \text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+} \), the former has the more negative electrode potential and so \( \text{Ag} \) will act as a reducing agent, reducing \( \text{Cr}_2\text{O}_7^{2-} \) to \( \text{Cr}^{3+} \):

\( \text{Ag}(s) \rightarrow \text{Ag}^+(aq) + 1 \text{e}^- \) and \( \text{Cr}_2\text{O}_7^{2-}(aq) + 14 \text{H}^+(aq) + 6 \text{e}^- \rightleftharpoons 2 \text{Cr}^{3+}(aq) + 7 \text{H}_2\text{O}(l) \)

Between \( \text{Ag}^+/\text{Ag} \) and \( \text{Cr}^{3+/\text{Cr}^{2+}} \), the latter has the more negative electrode potential and so no further reaction can occur i.e. the \( \text{Ag} \) cannot reduce any \( \text{Cr}^{3+} \) ions formed to \( \text{Cr}^{2+} \) ions.

Therefore the **final colour of the solution will be green** owing to the presence of \( \text{Cr}^{3+}(aq) \) ions.

The overall equation for the reaction is:

\[ 6 \text{Ag}(s) + \text{Cr}_2\text{O}_7^{2-}(aq) + 14 \text{H}^+(aq) \rightleftharpoons 6 \text{Ag}^+(aq) + 2 \text{Cr}^{3+}(aq) + 7 \text{H}_2\text{O}(l) \]

*(1 mark for the identification of the correct final colour, 1 mark for the overall equation)*
12. Redox equilibria answers

12.5. Applications of electrochemical cells

1.

<table>
<thead>
<tr>
<th>Battery type</th>
<th>Half equations with electrode potentials</th>
<th>$E_{\text{cell}}$ / V</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zinc/carbon</td>
<td>$\text{Zn}^{2+}(aq) + 2 \text{e}^- \rightleftharpoons \text{Zn}(s)$ [\text{E} \approx -0.8 \text{ V}] [\text{E} = +0.7 \text{ V}]</td>
<td>$\pm 1.5$</td>
</tr>
<tr>
<td></td>
<td>$2 \text{NH}_3(aq) + 2 \text{e}^- \rightleftharpoons 2 \text{NH}_3(g) + \text{H}_2(g)$</td>
<td></td>
</tr>
<tr>
<td>Nickel/cadmium</td>
<td>$\text{Cd(OH)}_2(s) + 2 \text{e}^- \rightleftharpoons \text{Cd}(s) + 2 \text{OH}^-\text{(aq)}$ [\text{E} \approx -0.8 \text{ V}]</td>
<td>$\pm 1.3$</td>
</tr>
<tr>
<td></td>
<td>$\text{NiO(OH)}(s) + \text{H}_2\text{O(l)} + \text{e}^- \rightleftharpoons \text{Ni(OH)}_2(s) + \text{OH}^-\text{(aq)}$ [\text{E} = +0.5 \text{ V}]</td>
<td></td>
</tr>
<tr>
<td>Lead-acid</td>
<td>$\text{PbSO}_4(s) + 2 \text{e}^- \rightleftharpoons \text{Pb}(s) + \text{SO}_4^{2-}\text{(aq)}$ [\text{E} \approx -0.35 \text{ V}]</td>
<td>$\pm 2.05$</td>
</tr>
<tr>
<td></td>
<td>$\text{PbO}_2(s) + 4 \text{H}^+(aq) + 2 \text{e}^- \rightleftharpoons \text{PbSO}_4(s) + 2 \text{H}_2\text{O(l)}$ [\text{E} = +1.70 \text{ V}]</td>
<td></td>
</tr>
<tr>
<td>Fuel cell</td>
<td>$2 \text{H}^+(aq) + 2 \text{e}^- \rightleftharpoons \text{H}_2(g)$ [\text{E} \approx 0.0 \text{ V}]</td>
<td>$\pm 1.2$</td>
</tr>
<tr>
<td></td>
<td>$4 \text{H}^+(aq) + \text{O}_2(g) + 4 \text{e}^- \rightleftharpoons 2 \text{H}_2\text{O(l)}$ [\text{E} = +1.2 \text{ V}]</td>
<td></td>
</tr>
</tbody>
</table>

2. (a) Nickel, from +3 in NiO(OH) to +2 in Ni(OH)$_2$ \( (4 \text{ marks}) \) \( (2 \text{ marks}) \)

(b) $[2 \text{OH}^-\text{(aq)} + \text{Cd(s)}], \text{Cd(OH)}_2(s) \parallel [\text{NiO(OH)}(s) + \text{H}_2\text{O(l)}], [\text{Ni(OH)}_2(s) + \text{OH}^-\text{(aq)}]$ \( (2 \text{ marks}, 1 \text{ for each side correct}) \)

(c) Cd(OH)$_2(s) + 2$ Ni(OH)$_2(s) \rightarrow 2$ NiO(OH)(s) + 2 H$_2$O(l) + Cd(s) \( (1 \text{ mark}) \)

(d) Nickel / cadmium are both toxic metals. \( (1 \text{ mark}) \)