National Certificate of Educational Achievement
TAUMATA MĀTAURANGA Ā-MOTU KUA TAEA

Exemplar for Internal Achievement Standard

Chemistry Level 3

TRIAL

This exemplar supports assessment against:

Achievement Standard AS91393

Demonstrate understanding of oxidation-reduction processes

An annotated exemplar is an extract of student evidence, with a commentary, to explain key aspects of the standard. It assists teachers to make assessment judgements at the grade boundaries.

New Zealand Qualifications Authority

To support internal assessment

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	Grade: Excellence
	For Excellence, the student needs to demonstrate understanding of oxidation-reduction processes. This involves producing evidence that shows justification of the spontaneity of electrochemical and electrolytic cells, supported by equations and calculations. The student needs to produce evidence that links the balanced half and full equations and either cell potential calculations or using reduction potentials to justify the spontaneity of electrochemical and electrolytic cells. This student has explained oxidation and reduction with supporting half and full equations (1). The student has also calculated the cell potential for both cells (2), and used the reduction potentials to compare the relative oxidant/reductant strength of the REDOX couples (3) for both electrochemical and electrolytic cells. The student has provided evidence of using all their supporting evidence to justify the spontaneity (4) of both electrochemical and electrolytic cells.

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Student 1: Excellence

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Electrochemical cell:

Reduction: $Cu^{2+} + 2e^- \rightarrow Cu$

Oxidation: $Zn \rightarrow Zn^{2+} + 2e^-$ (1)

Overall: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

Cu^{2+} is reduced due to the fact that it gains 2 electrons in order to become Cu and reduction is the gain of electrons. The oxidation number for Cu decreases from +2 in $Cu^{2+}$to 0 in Cu which shows that this is a reduction reaction.

Zn is oxidised due to the fact that it loses 2 electrons in order to become Zn^{2+} and oxidation is the loss of electrons. The oxidation number for Zn increases from 0 in Zn to +2 in Zn^{2+} which shows that this is an oxidation reaction.

At the right-hand electrode, the blue Cu^{2+} solution is reduced to form a brown Cu solid which is deposited on the brown Cu electrode causing it to gain mass. The blue solution in the beaker will lighten as Cu^{2+} is reduced.

At the left-hand electrode, the silver Zn solid is oxidised forming the colourless solution of Zn^{2+}. This means that the silver/grey Zn electrode will decrease in mass. The solution in the beaker remains colourless.

$E°_{cell} = E°reduction – E°oxidation$

$=+0.34 – -0.76$ (2)

$= 1.10\ V$

Since the $E°_{cell}$ is positive this means that the reaction is spontaneous in the direction shown. Therefore, electrochemical potential energy is converted into electrical energy which is seen through the 1.10 V of electrical energy produced by the cell under standard conditions.

$E°_{cell}\ (Zn^{2+}/Zn) = -0.76V$ $E°_{cell}\ (Cu^{2+}/Cu) = +0.34V$

$(Cu^{2+}/Cu) > (Zn^{2+}/Zn)$ Therefore Cu^{2+} is reduced because it has a more positive $E°_{cell}$ and the right-hand electrode gains mass and Zn oxidised because it has a more negative $E°_{cell}$ and the left-hand electrode losses mass. Also, because the Cu^{2+} is has a more positive $E°_{cell}$ it means that it is the stronger oxidant and therefore the reactions will occur spontaneously. (3)

Electrolytic cell:

Since both the graphite electrodes are inert, then either Cu^{2+} or H_2O is reduced and SO_4^{2-} or H_2O is oxidised.

Since the right-hand electrode gains mass then Cu^{2+} must be reduced.

Since bubbles which relight a glowing splint are observed at the left-hand electrode, this indicates the presence of O_2, meaning water must be oxidised. As well as this SO_4^{2-} is fully oxidised already.

Reduction: $Cu^{2+} + 2e^- \rightarrow Cu$

Oxidation: $2H_2O \rightarrow O^{2-} + H^+ + 4e^-$ (1)

Overall: $2Cu^{2+} + 2H_2O \rightarrow 2Cu + O^{2-} + H^+$

Cu^{2+} is reduced because it gains 2 electrons in order to become Cu and reduction is the gain of electrons. The oxidation number for Cu decreases from +2 in $Cu^{2+}$to 0 in Cu which also shows that this is a reduction reaction.

$2H_2O$ is oxidised because it loses 4 electrons in order to become O_2 and oxidation is the loss of electrons. The oxidation number for O increases from -2 in H_2O to 0 in O_2 which shows that this is an oxidation reaction.

At the right-hand electrode, the blue Cu^{2+} solution is reduced to brown Cu solid which is deposited on the right-hand electrode (cathode) which therefore gain mass. The blue solution in the beaker will lighten as Cu^{2+} is reduced. At the left-hand electrode, colourless H_2O liquid

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is oxidised to the colourless O_2 gas, which is released as bubbles at the anode and these bubbles can relight a glowing splint.

$E°_{cell}$ = E°reduction – E°oxidation

=+0.34 – + 1.23

= -0.89 V

2

Since the $E°_{cell}$ is negative, this means that the reaction is non-spontaneous in the direction shown. This means that in order for the reaction to occur, an input of electrical energy of at least 0.90 V is required from the power supply.

3

(Cu^{2+}/Cu) < (O_2/H_2O) Cu^{2+} has a more negative $E°_{cell}$ and H_2O is more positive. This means that as the reaction is shown it will not proceed because Cu^{2+} is a stronger reductant than H_2O. So, because Cu^{2+} is has a more negative $E°_{cell}$ value it means that the reaction will not occur spontaneously.

Both the electrochemical and electrolytic cell reactions produce the brown Cu metal at the right-hand electrode through the reduction of Cu^{2+}/Cu. However, the way this is achieved is different. The REDOX reaction in the electrochemical cell which forms Cu is a spontaneous reaction where electrochemical potential energy is converted into electrical energy. This reaction produces 1.10 V under standard conditions. In comparison, the REDOX reaction in the electrolytic cell where Cu is produced is a non-spontaneous reaction, and requires the input of electrical energy to force the reaction to occur. The electrolytic cell requires more than 0.89 V of energy to make the reaction proceed. Therefore, electrochemical cells convert electrochemical potential energy into electrical energy to produce Cu, whereas electrolytic cells require electrical energy to produce Cu.

4

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	Grade Boundary: High Merit
	For Merit, the student needs to undertake effective project management to support technological practice. This involves making and explaining links between oxidation-reduction processes, observations, equations, and calculations. This requires explanations that use chemistry vocabulary, symbols, and conventions. This student has identified what has been oxidised and reduced with reasons (1), and given balanced half equations for electrochemical and electrolytic cells (2). The student has also completed a correct reduction potential calculation for the electrochemical cell (3), explained the spontaneity of the reaction (4), and related species to the given observations (5). This student begins to link the spontaneity of the electrochemical cell to the cell potential calculation. The student also includes balanced full equations. This evidence does not reach Excellence because the student needs to justify the spontaneity of both electrochemical and electrolytic cells by using the cell potential calculation(s) or the standard reduction potentials.

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Student 2: High Merit

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Electrochemical Cell

Reduction at the cathode:

(2) $H_2O_2 + 2H^+ + 2e \rightarrow 2H_2O$

The solution remains colourless as H_2O_2 and H_2O are both colourless.

This is a reduction reaction as the oxidation number of O decreases from -1 in H_2O_2 to -2 in H_2O. A decrease in oxidation number corresponds to reduction.

Each H_2O_2 gains two electrons. Gaining of electrons corresponds to reduction.

Oxidation at the anode:

(1) $Fe^{2+} \rightarrow Fe^{3+} + e$

The solution turns from pale green Fe^{2+} to orange Fe^{3+}.

This is an oxidation reaction as the oxidation number of Fe increases from +2 in Fe^{2+} to +3 in Fe^{3+}. An increase in oxidation number corresponds to oxidation.

Each Fe^{2+} loses 1 electron, losing electrons corresponds to oxidation.

E°

An electrochemical cell is an apparatus that uses a spontaneous oxidation-reduction reaction to produce an electric current. It consists of two half cells connected by a conducting wire and a salt bridge. The reaction is spontaneous creating electrical energy.

For this reaction:

(3) $E^o cell = E^o red - E^o ox = +1.77\ V - 0.77\ V = +1.00\ V$

(4) As $E^o cell$ is positive, the reaction is spontaneous.

The most positive E^o value will be the reduction reaction as these are reduction potentials

Overall equation:

$2Fe^{2+}(aq) + H_2O_2(aq) + H^+(aq) \rightarrow 2H_2O(l) + 2Fe^{3+}(aq)$

Electolytic cell

From observations: The electrolysis of molten copper oxide produces bubbles of a colourless gas at one electrode and an orange solid is deposited at the other electrode.

(5) Colourless gas is oxygen at the positive anode. Orange solid is copper and it is deposited at the negative cathode.

Reduction at the cathode (negative electrode)

$Cu^{2+} + 2e \rightarrow Cu$

(2) The oxidation number of Cu decreases from +2 in Cu^{2+} to 0 in Cu.

Each Cu^{2+} gains 2 electrons.

Oxidation at the anode (positive electrode)

(1) $2O^{2-} \rightarrow O_2 + 4e$

The bubbles of colourless gas are oxygen.

The oxidation number of O increases from -2 in O^{2-} to 0 in O_2.

Each O^{2-} loses 2 electrons.

Overall equation

$2Cu^{2+}(l) + 2O^{2-}(l) \rightarrow 2Cu(s) + O_2(g)$

Energy requirements

The battery provides push for electrons to move in the reverse direction to their natural tendency. The non-spontaneous reaction is forced to occur and electrical energy is consumed in order to produce a chemical reaction.

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	Grade Boundary: Low Merit
3.	For Merit, the student needs to demonstrate in-depth understanding of oxidation-reduction processes. This involves making and explaining links between oxidation-reduction processes, observations, equations and calculations. This requires explanations that use chemistry vocabulary, symbols, and conventions. This student has identified what has been oxidised and reduced with reasons (1) and given balanced half equations for electrochemical and electrolytic cells (2). The student has also completed a reduction potential calculation for the electrochemical cell (3), and related some species to given observations (4). For a more secure Merit, the student could relate all species to given observations, and included the correct unit for the calculation.

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Student 3: Low Merit

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Electrochemical Cell

Reduction at the cathode:

(2) $H_2O_2 + 2H^+ + e \rightarrow 2H_2O$

The reaction remains colourless.

(1) This is a reduction reaction as the oxidation number of O decreases from -1 in H_2O_2 to -2 in H_2O. A decrease in oxidation number corresponds to reduction.

Oxidation at the anode:

$Fe^{2+} \rightarrow Fe^{3+} + e$

The solution turns from pale green Fe^{2+} to orange Fe^{3+}.

This is an oxidation reaction as the oxidation number of Fe increases from +2 in Fe^{2+} to +3 in Fe^{3+}. An increase in oxidation number corresponds to oxidation.

Each Fe^{2+} loses 1 electron, losing electrons corresponds to oxidation.

E^o

For this reaction:

(3) $E^o cell = E^o red - E^o ox = +1.77 - 0.77 = +1.00$

The reaction is spontaneous.

Electolytic cell

From observations: The electrolysis of molten copper oxide produces bubbles of a colourless

(4) gas at one electrode and an orange solid is deposited at the other electrode.

Colourless gas is oxygen at the positive anode. Orange solid is copper and it is deposited at the negative cathode.

Reduction at the cathode (negative electrode)

(2) $Cu^{2+} + 2e \rightarrow Cu$

The oxidation number of Cu decreases from +2 in Cu^{2+} to 0 in Cu.

Each Cu^{2+} gains 2 electrons.

Oxidation at the anode (positive electrode)

$2O^{2-} \rightarrow O_2 + 4e$

The bubbles of colourless gas are oxygen.

Each O^{2-} loses 2 electrons.

This process requires electrical energy to make the non-spontaneous reaction happen.

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	Grade Boundary: High Achieved
4.	For Achieved, the student needs to demonstrate understanding of oxidation-reduction processes. This involves describing oxidation-reduction processes and may involve calculations. This requires the use of chemistry vocabulary, symbols, and conventions. This student has identified what has been oxidised and reduced for electrochemical and electrolytic cells, with a description of either loss/gain of electrons or oxidation number (ON) changes (1). They have also made a reference to the requirement of energy for electrolytic cell (2) and reduction potentials for electrochemical cell (3). To reach Merit, the student could include correct half equations for all reactions.

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Student 4: High Achieved

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Electrochemical Cell

Reduction at the cathode:

$H_2O_2 + e \rightarrow 2H_2O$

The reaction remains colourless.

This is a reduction reaction as the oxidation number of O decreases from -1 in H_2O_2 to -2 in H_2O. A decrease in oxidation number corresponds to reduction. (1)

Oxidation at the anode:

$Fe^{2+} \rightarrow Fe^{3+} + e$

The solution turns from pale green Fe^{2+} to orange Fe^{3+}.

This is an oxidation reaction as the oxidation number of Fe increases from +2 in Fe^{2+} to +3 in Fe^{3+}. An increase in oxidation number corresponds to oxidation.

Each Fe^{2+} loses 1 electron, losing electrons corresponds to oxidation.

E^o

For this reaction:

E^0 for H_2O_2 is the most positive (1.77) so it will be reduced. (3)

The reaction is spontaneous creating electrical energy.

Electolytic cell

From observations: The electrolysis of molten copper oxide produces bubbles of a colourless gas at one electrode and an orange solid is deposited at the other electrode.

Colourless gas is oxygen at the positive anode. Orange solid is copper and it is deposited at the negative cathode.

Reduction at the cathode (negative electrode)

$Cu^{2+} + 2e \rightarrow Cu$

The oxidation number of Cu decreases from +2 in Cu^{2+} to 0 in Cu.

Each Cu^{2+} gains 2 electrons.

Oxidation at the anode (positive electrode)

$O^{2-} \rightarrow O_2 + 4e$

The bubbles of colourless gas are oxygen.

Each O^{2-} loses 2 electrons. (2)

This process requires electrical energy.

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	Grade Boundary: Low Achieved
5.	For Achieved, the student needs to demonstrate understanding of oxidation-reduction processes. This involves describing oxidation-reduction processes and may involve calculations. This requires the use of chemistry vocabulary, symbols, and conventions. This student has identified what has been oxidised and reduced for electrochemical and electrolytic cells, with descriptions of either loss/gain of electrons or oxidation number (ON) changes (1), some reference to the requirement of energy for the electrolytic cell (2), and the reduction potentials for the electrochemical cell (3). For a more secure Achieved, the student could link oxidation number increase/decrease or loss/gain of electrons to oxidation and reduction processes.

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Student 5: Low Achieved

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Electrochemical Cell

Reduction at the cathode:

$H_2O_2 \rightarrow 2H_2O$

The reaction remains colourless.

1 This is a reduction reaction as the oxidation number of O changes from -1 in H_2O_2 to -2 in H_2O

Oxidation at the anode:

$Fe^{2+} \rightarrow Fe^{3+} + e$

This is an oxidation reaction as the oxidation number of Fe increases from +2 in Fe^{2+} to +3 in Fe^{3+}. An increase in oxidation number corresponds to oxidation.

Each Fe^{2+} loses 1 electron, losing electrons corresponds to oxidation.

Eº

3 E^0 for H_2O_2 is the most positive so it will be reduced. This reaction creates electrical energy.

Electolytic cell

From observations: The electrolysis of molten copper oxide produces bubbles of a colourless gas at one electrode and an orange solid is deposited at the other electrode.

Colourless gas is oxygen at the positive anode. Orange solid is copper and it is deposited at the negative cathode.

Reduction at the cathode (negative electrode)

$Cu^{2+} \rightarrow Cu$

The oxidation number of Cu goes from +2 in Cu^{2+} to 0 in Cu.

Each Cu^{2+} gains 2 electrons.

Oxidation at the anode (positive electrode)

$O^{2-} \rightarrow O_2 + 4e$

The bubbles of colourless gas are oxygen.

Each O^{2-} loses 2 electrons.

2 This process requires external electrical energy.

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	Grade Boundary: High Not Achieved
6.	For Achieved, the student needs to demonstrate understanding of oxidation-reduction processes. This involves describing oxidation-reduction processes and may involve calculations. This requires the use of chemistry vocabulary, symbols, and conventions. This student has identified what has been oxidised and reduced for electrochemical and electrolytic cells, with some reference to either loss/gain of electrons or oxidation number (ON) changes (1). To reach Achieved, the student could ensure that they have described all oxidation and reduction changes correctly, and made a reference to the requirement of energy for the electrolytic cell.

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Student 6: High Not Achieved

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Electrochemical Cell

Reduction at the cathode:

$H_2O_2 \rightarrow 2H_2O$

This is a reduction reaction as the oxidation number of O changes from 0 in H_2O_2 to -2 in H_2O

1

Oxidation at the anode:

$Fe^{2+} \rightarrow Fe^{3+} + e$

This is an oxidation reaction as the oxidation number of Fe increases from +2 in Fe^{2+} to +3 in Fe^{3+}. An increase in oxidation number corresponds to oxidation.

Each Fe^{2+} loses 1 electron.

E°

E^0 for H_2O_2 is the most positive so it will be reduced. This reaction creates electrical energy.

Electolytic cell

From observations: The electrolysis of molten copper oxide produces bubbles of a colourless gas at one electrode and an orange solid is deposited at the other electrode. Colourless gas is oxygen at the positive anode. Orange solid is copper and it is deposited at the negative cathode.

Reduction at the cathode (negative electrode)

$Cu^{2+} \rightarrow Cu$

The oxidation number of Cu goes from +2 in Cu^{2+} to 0 in Cu.

Oxidation at the anode (positive electrode)

$O^{2-} \rightarrow O_2 + 4e$

The bubbles of colourless gas are oxygen.

Each O^{2-} loses 2 electrons.

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