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90311

NEW ZEALAND QUALIFICATIONS AUTHORITY
MANA TOHU MĀTAURANGA O AOTEAROA

National Certificate of Educational Achievement
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Level 2 Chemistry, 2006

90311 Describe oxidation–reduction reactions

Credits: Three
2.00 pm Monday 27 November 2006

Check that the National Student Number (NSN) on your admission slip is the same as the number at the top of this page.

A Periodic Table is provided on the RESOURCE SHEET in your Level 2 Chemistry package.

You should answer ALL the questions in this booklet.

If you need more space for any answer, use the page(s) provided at the back of this booklet and clearly number the question.

Check that this booklet has pages 2–8 in the correct order and that none of these pages is blank.

YOU MUST HAND THIS BOOKLET TO THE SUPERVISOR AT THE END OF THE EXAMINATION.

For Assessor's use only	**Achievement Criteria**				
Achievement		**Achievement with Merit**		**Achievement with Excellence**	
Describe oxidation-reduction reactions.		Apply oxidation-reduction principles.		Discuss oxidation-reduction processes.	
Overall Level of Performance					

You are advised to spend 35 minutes answering the questions in this booklet.

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QUESTION ONE

(a) The following two test tube reactions were carried out in the laboratory.

Identify the species responsible for each of the observations.

(i) Freshly made iron(II) sulfate was added drop-wise to acidified potassium permanganate solution. The solution turned from **purple** to **colourless**.

Identify the ion responsible for the **purple** colour _____

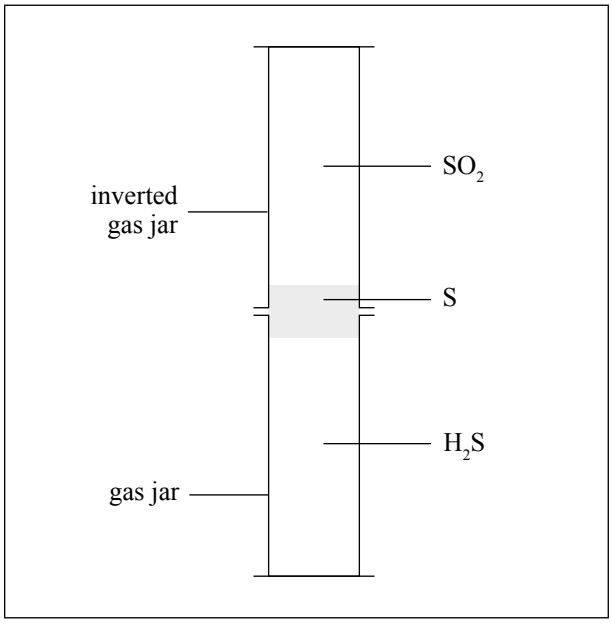
Identify the ion formed as the purple colour goes **colourless** _____

(ii) Chlorine water was added to potassium iodide solution in a test tube. Eventually a grey solid was deposited in the test tube.

Identify the grey solid _____

(b) A gas jar was filled with hydrogen sulfide gas, $H_2S(g)$, and a second gas jar was filled with sulfur dioxide gas, $SO_2(g)$.

The gas jar of SO_2 was inverted over the gas jar of H_2S as shown in the diagram opposite. After a short time, yellow sulfur, S, was seen where the two gases met.



An unbalanced equation for the reaction between H_2S and SO_2 can be represented as:

$$H_2S + SO_2 \rightarrow S + H_2O$$

(i) Identify the **oxidant** and **reductant** in this reaction.

Oxidant: _____

Reductant: _____

(ii) Write balanced half-equations for the oxidation and reduction reactions.

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Oxidation

Reduction

(iii) Combine these two half-equations to give the balanced equation for this redox reaction.

QUESTION TWO

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The box below shows a reaction scheme for some compounds of chromium.

Reaction Scheme

$Cr_2O_3(s)$	→	$CrO_4^{2-}(aq)$	→	$Cr_2O_7^{2-}(aq)$	→	$Cr^{3+}(aq)$

(a) Determine the oxidation number of chromium in each of these four chromium species.

Cr_2O_3 _____

CrO_4^{2-} _____

$Cr_2O_7^{2-}$ _____

Cr^{3+} _____

(b) Circle **ONE** step in this scheme that shows the chromium species being **reduced**.

$Cr_2O_3 \rightarrow CrO_4^{2-}$ $\quad$ $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$ $\quad$ $Cr_2O_7^{2-} \rightarrow Cr^{3+}$

Justify your answer.

_____

(c) Complete the balanced half-equation for the reaction $Cr_2O_3 \rightarrow CrO_4^{2-}$

Cr_2O_3	→	CrO_4^{2-}

(d) (i) Complete the half-equation for the conversion $Cr_2O_7^{2-} \rightarrow Cr^{3+}$

$Cr_2O_7^{2-}$	→	Cr^{3+}

(ii) Describe the expected observations that occur as the conversion $Cr_2O_7^{2-} \rightarrow Cr^{3+}$ proceeds.

Link these observations to the species involved in the reaction.

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QUESTION THREE

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(a) Many elements exhibit variable oxidation numbers. The oxidation numbers of copper are 0, +1 and +2.

The element copper is present in many metallic minerals. These minerals include chalcopyrite ($CuFeS_2$), malachite ($Cu_2(CO_3)(OH)_2$), and copper metal (Cu).

(i) State the oxidation number of copper in copper metal (Cu) _____

(ii) Determine the oxidation number of copper in malachite, $Cu_2(CO_3)(OH)_2$. Include your working to show how the oxidation number was determined.

_____

_____

(iii) Sulfur has an oxidation number of –2 in chalcopyrite, $CuFeS_2$.

Discuss the possible oxidation numbers of both copper and iron in chalcopyrite, $CuFeS_2$.

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(b) Copper ores that contained the compound copper(II) sulfide, CuS, were first mined by ancient miners. First they roasted the CuS in air as shown by **Equation A**. Then the product was reacted with carbon as shown by **Equation B**.

Roasting in air: $2CuS + 3O_2 \rightarrow 2CuO + 2SO_2$ **Equation A**

Reacting the product with carbon: $CuO + C \rightarrow Cu + CO$ **Equation B**

Both these reactions are examples of redox equations.

(i) In Equation A circle the **element** which is **oxidised**.

(ii) Justify your answer.

_____

_____

(iii) In terms of oxidation number, explain why **Equation B** is a redox reaction.

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QUESTION FOUR

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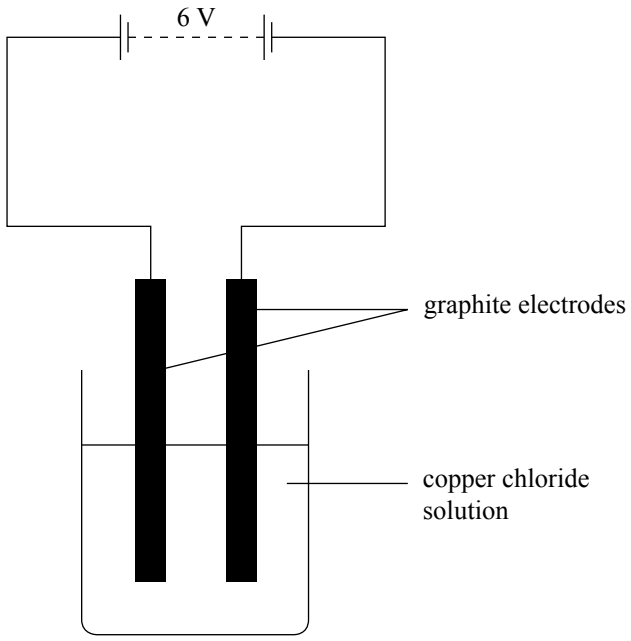
The diagram below shows the electrolysis of an aqueous solution of copper chloride.



During this electrolysis, a student noticed that copper metal was deposited at the cathode and chlorine gas was given off at the anode.

(a) (i) Write the half-equation for the reaction that occurs at the cathode.

(ii) Write the half-equation for the reaction that occurs at the anode.

(b) Write a paragraph discussing how copper is produced at the cathode and chlorine is produced at the anode.

Use your knowledge of the principles of electrolysis and refer to the diagram opposite. The the key words below should be used in your answer.

Key words:
anode, cathode, electrons, anions, cations, negative, positive, oxidation, reduction

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Extra paper for continuation of answers if required.
Clearly number the question.

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Question number	