SACE 1 Chemistry   **Redox & Electrochemistry Test 2018**   *ANSWERS* Name

1. Oxygen is the most electronegative **common** element, with only fluorine having a higher value. Determine the oxidation state for oxygen in the following substances.   [5 marks]

Water H_2O	Hydrogen Peroxide H_2O_2	Oxygen Difluoride F_2O	Ozone O_3	Thionyl Fluoride SOF_2
oxidation state *−2*	oxidation state *−1*	oxidation state *+2*	oxidation state *0*	oxidation state *−2*

2. Oxidation state represents the number of chemical bonds an atom can make with different elements in a compound.

   (a) State the three rules arising from this definition of Oxidation State, including significant variations when appropriate: [6 marks]

	State the rule	Appropriate variations
Rule 1 for elements	*All elements have oxidation state zero, 0*	*Monatomic ion charge*
Rule 2 for oxygen	*Oxygen usually has oxidation state −2*	*• −1 with Peroxides* *• +2 with Fluorine*
Rule 3 for hydrogen	*Hydrogen has oxidation state +1 with non-metals*	*−1 when with metals and most metalloids*

   (b) Explain how the Oxidation State for all other elements are determined using these rules:

   Assign oxidation state for any elemental species, oxygen and hydrogen first, and THEN adjust OS of any remaining atoms so that total OS adds up to the overall ionic charge.   [2marks]

3. Determine oxidation state of the following **highlighted** element **P**hosphorous: [4 marks]

Phosphine PH_3	Phosphoric Acid H_3PO_4	Phosphorous Acid H_3PO_3	Phosphorous Hydride P_2H_4
oxidation state **P** *−3*	oxidation state **P** *+5*	oxidation state **P** *+3*	oxidation state **P** *−2*

4. State which of these reactions are Redox by including the Oxidation State of the **underlined** elements. [5 marks]
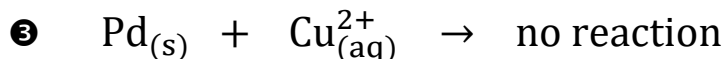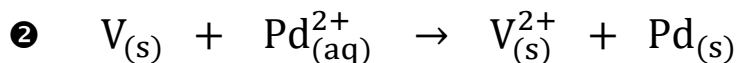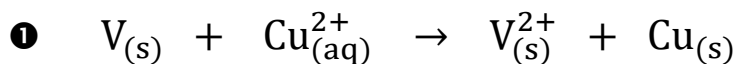
(a)	$$2H_2 \overset{-1}{\underline{O}}_2 \rightarrow \overset{0}{\underline{O}}_2 + 2H_2 \overset{-2}{\underline{O}}$$	*Redox*
(b)	$$K_2 \overset{+6}{\underline{Cr}}_2 O_7 + 2K\overset{-2}{\underline{O}}H \rightarrow 2K_2 \overset{+6}{\underline{Cr}}O_4 + H_2 \overset{-2}{\underline{O}}$$	*Not Redox*
(c)	$$\overset{0}{\underline{Cu}} + 4H\overset{+5}{\underline{N}}O_3 \rightarrow \overset{+2}{\underline{Cu}}(\overset{+5}{\underline{N}}O_3)_2 + 2\overset{+4}{\underline{N}}O_2 + 2H_2O$$	*Redox*
(d)	$$Na_2 \overset{+2}{\underline{S}}_2 O_3 + 2H\overset{-1}{\underline{Cl}} \rightarrow \overset{+4}{\underline{S}}O_2 + \overset{0}{\underline{S}} + H_2O + 2Na\overset{-1}{\underline{Cl}}$$	*Redox*
(e)	$$\overset{+6}{\underline{Cr}}_2 O_7^{2-} + 8HCl + 3H_2 \overset{-1}{\underline{O}}_2 \rightarrow 2\overset{+3}{\underline{Cr}}Cl_3 + 3\overset{0}{\underline{O}}_2 + 7H_2O + 2Cl^{1-}$$	*Redox*

5. Complete the Redox reactions using the redox conjugate pairs provided. [2 each = 10 marks]

	Conjugate pairs	Using ½-equation method to balance full redox reaction equation
a)	$Ca \mid Ca^{2+}$ with $H_2O \mid H_2$	$$Ca \rightarrow Ca^{2+} + 2e^-$$ $$2H_2O + 2e^- \rightarrow H_2 + 2OH^{1-}$$ ___ $$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$
b)	$PbO \mid Pb$ with $Br_2 \mid BrO_3^{1-}$	$$[\ PbO + 2H^+ + 2e^- \rightarrow Pb + H_2O\ ] \times 5$$ $$Br_2 + 6H_2O \rightarrow 2BrO_3^{1-} + 12H^+ + 10e^-$$ ___ $$5PbO + H_2O + Br_2 \rightarrow 5Pb + 2H^+ + 2BrO_3^{1-}$$
c)	$Cr_2O_7^{2-} \mid Cr^{3+}$ with $Fe^{2+} \mid Fe^{3+}$	$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$ $$[\ Fe^{2+} \rightarrow Fe^{3+} + e^-\ ] \times 6$$ ___ $$Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow 2Cr^{3+} + 7H_2O + 6Fe^{3+}$$
d)	$MnO_2 \mid Mn_2O_3$ with $Fe^{2+} \mid Fe^{3+}$	$$2MnO_2 + 2H^+ + 2e^- \rightarrow Mn_2O_3 + H_2O$$ $$[\ Fe^{2+} \rightarrow Fe^{3+} + e^-\ ] \times 2$$ ___ $$2MnO_2 + 2H^+ + 2Fe^{2+} \rightarrow Mn_2O_3 + H_2O + 2Fe^{3+}$$
e)	$C_2H_5OH \mid CH_3COOH$ with $MnO_4^{1-} \mid Mn^{2+}$	$$[\ C_2H_5OH + H_2O \rightarrow CH_3COOH + 4H^+ + 4e^-\ ] \times 5$$ $$[\ MnO_4^{1-} + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O\ ] \times 4$$ ___ $$5C_2H_5OH + 12H^+ + 4MnO_4^{1-} \rightarrow 5CH_3COOH + 11H_2O + 4Mn^{2+}$$

6. Copper (Cu), palladium (Pd) and vanadium (V) are metals used in making alloys. Three displacement tests were carried out on these three metals $M_{(s)}$ and their corresponding salts $M^{n+}_{(aq)}$ in solution. The results are summarised below:

❶ $V_{(s)} + Cu^{2+}_{(aq)} \rightarrow V^{2+}_{(s)} + Cu_{(s)}$

❷ $V_{(s)} + Pd^{2+}_{(aq)} \rightarrow V^{2+}_{(s)} + Pd_{(s)}$

❸ $Pd_{(s)} + Cu^{2+}_{(aq)} \rightarrow$ no reaction

(a) (i) Complete the displacement reaction equation for the following combination:

❹ $Cu_{(s)} + Pd^{2+}_{(aq)} \rightarrow$ **[1 mark]**

(i) Arrange the three metals in order of **increasing** reactivity: **[2 marks]**

Pd < Cu < V

(ii) Provide a brief explanation for your answer in part **(a) (i)** above. **[3 marks]**

- Reactions ❶ and ❷ show that both Cu and Pd are displaced by V, so that reactivity is $\frac{Cu}{Pd} < V$.
- ❸ shows that Cu is not displaced by Pd, so that $Pd < Cu$.
- Overall, $Pd < Cu < V$.
  (Alternatively, electronegativity values $^{Pd}_{2.20}\chi > ^{Cu}_{1.90}\chi > ^{V}_{1.63}\chi$)

(b) Two half-cells for $V\,|\,V^{2+}$ and $Pd\,|\,Pd^{2+}$ are connected with a salt bridge soaked in potassium nitrate solution and a multimeter set to record DC voltages as shown opposite.
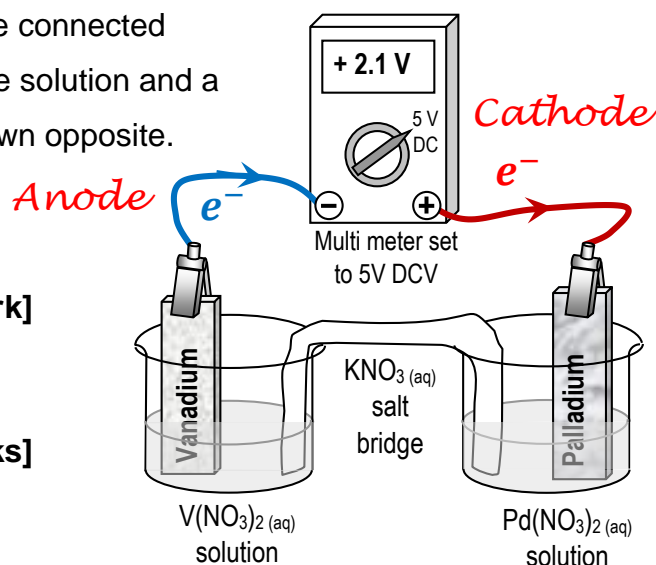
(i) Indicate the direction of the external electron flow with an arrow on the diagram opposite. **[1 mark]**

(ii) Complete the table below by referring to the (+) and (−) electrodes shown in the diagram opposite. **[4 marks]**

+ 2.1 V
5 V DC
Anode e^-
Cathode e^-
Multi meter set to 5V DCV
Vanadium
KNO$_3$ (aq) salt bridge
Palladium
V(NO$_3$)$_2$ (aq) solution
Pd(NO$_3$)$_2$ (aq) solution

Electrode Sign		½-cell M | M^{n+} at electrode	Redox Process Name (Reduction or oxidation)	Redox Process Name (Reduction or Oxidation)	Electrode Name (Cathode or Anode?)
From Diagram	(+)	$Pd^{2+} \mid Pd$	$Pd^{2+} + 2e^- \rightarrow Pd$	*Reduction*	*Cathode*
	(−)	$V \mid V^{2+}$	$V \rightarrow V^{2+} + 2e^-$	*Oxidation*	*Anode*

(iii)	Write full redox equation for these ½-reactions	$V + Pd^{2+} \rightarrow V^{2+} + Pd$	Which displacement reaction ❶, ❷, ❸ or ❹ is represented by connecting these two ½-cells? #❷ **[2 marks]**

(c) (i) The $V\,|\,V^{2+}$ half-cell has a solution containing $V^{2+}_{(aq)}$ and $NO^{1-}_{3(aq)}$ ions.

Which one of these ions becomes in excess in the solution as the redox reaction proceeds? **[1 mark]**

$V \rightarrow V^{2+} + 2e^-$ *produces more* V^{2+} *to enter the solution.*
This makes V^{2+} *in excess compared to* NO^{1-}_3

(ii) The $Pd\,|\,Pd^{2+}$ half-cell has a solution containing $Pd^{2+}_{(aq)}$ and $NO^{1-}_{3(aq)}$ ions.

Which one of these ions becomes in excess in the solution as the redox reaction proceeds? **[1 mark]**

$Pd^{2+} + 2e^- \rightarrow Pd$ *removes* Pd^{2+} *from the solution.*
This makes NO^{1-}_3 *in excess compared to* Pd^{2+}.

(iii) The salt bridge contains mobile $K^{1+}_{(aq)}$ and $NO^{1-}_{3(aq)}$ ions.

State which half-cell each of these mobile ions move to as the redox reaction proceeds. **[2 marks]**

$K^{1+}_{(aq)}$ *moves to* NO^{1-}_3 *excess at* $\underbrace{Pd^{2+}|Pd}_{(+)\,cathode}$ *and* $NO^{1-}_{3(aq)}$ *moves to* V^{2+} *excess at* $\underbrace{V\,|\,V^{2+}}_{(-)\,anode}$

(d) The vanadium redox battery (VRB) uses vanadium salt solutions for both half cells:

$$VO^{1+}_2\,|\,VO^{2+}\,\|\,V^{2+}\,|\,V^{3+}$$
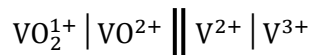
These are separated by a plastic membrane ($\|$) that allows aqueous ions to move between each half cell.

(i) State the four oxidation states of vanadium in the VRB: $\overset{+5}{V}O^{1+}_2\,|\,\overset{+4}{V}O^{2+}\,\|\,\overset{+2}{V}{}^{2+}\,|\,\overset{+3}{V}{}^{3+}$

[3 marks]

(ii) The VRB is able to convert chemical energy to electrical energy as an electrochemical cell and can also use electrical energy to produce a chemical reaction as an electrolytic cell. Refer to the two diagrams below and identify which process is occurring in each. **[2 marks]**
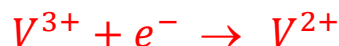
(iii) Indicate the direction of external electron flow on each diagram. **[2 marks]**

Electrochemical   *Electrolytic*

$VO^{1+}_{2(aq)}\,|\,VO^{2+}_{(aq)}\,\|\,V^{2+}_{(aq)}\,|\,V^{3+}_{(aq)}$

$VO^{2+}_{(aq)}\,|\,VO^{1+}_{2(aq)}\,\|\,V^{3+}_{(aq)}\,|\,V^{2+}_{(aq)}$

(iv) Use the half equation method to write the full redox reaction equation for the electrolytic VRB shown above.

[3 marks]

½ equation $V^{3+} + e^- \rightarrow V^{2+}$

½ equation $VO^{2+} + H_2O \rightarrow VO^{1+}_2 + 2H^{1+} + e^-$

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full equation $VO^{2+} + H_2O + V^{3+} \rightarrow VO^{1+}_2 + 2H^{1+} + V^{2+}$