

# ELECTROCHEMICAL CELLS

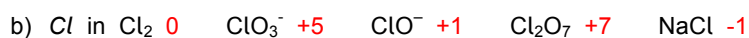
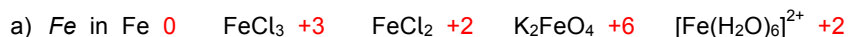


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## TASK 1 – Oxidation states

1) Calculate the oxidation state of the stated element in the following species:

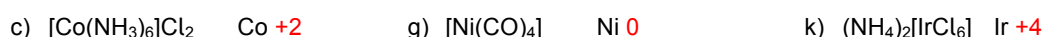
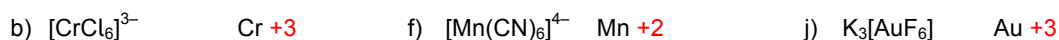
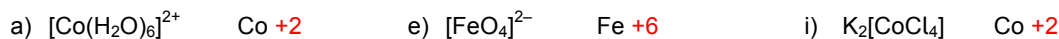


2) Calculate the oxidation state of each element in the following:

SO <sub>2</sub>	S +4	O -2
S	S 0	
SO <sub>3</sub>	S +6	O -2
H <sub>2</sub> S	H +1	S -2
NH <sub>3</sub>	N -3	H +1
NO <sub>2</sub>	N +4	O -2
N <sub>2</sub>	N 0	
NO <sub>3</sub> <sup>-</sup>	N +5	O -2
Cl <sup>-</sup>	Cl -1	
SO <sub>4</sub> <sup>2-</sup>	S +6	O -2

H <sub>2</sub> O <sub>2</sub>	H +1	O -1	
NaH	Na +1	H -1	
CO <sub>3</sub> <sup>2-</sup>	C +4	O -2	
Cr <sub>2</sub> O <sub>3</sub>	Cr +3	O -2	
CrO <sub>3</sub>	Cr +6	O -2	
KMnO <sub>4</sub>	K +1	Mn +7	O -2
K <sub>2</sub> MnO <sub>4</sub>	K +1	Mn +6	O -2
Cu <sub>2</sub> O	Cu +1	O -2	
CuO	Cu +2	O -2	
NaCuCl <sub>2</sub>	Na +1	Cu +1	Cl -1

3) Give the oxidation state of the transition metal in each of the following species.



## TASK 2 – Writing half equations



### TASK 3 – Combining half equations

- 1)  $\text{Zn} + 2\text{Fe}^{3+} \rightarrow \text{Zn}^{2+} + 2\text{Fe}^{2+}$
- 2)  $\text{Cr}_2\text{O}_7^{2-} + 6\text{Fe}^{2+} + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} + 6\text{Fe}^{3+}$
- 3)  $\text{H}_2\text{SO}_4 + 8\text{I}^- + 8\text{H}^+ \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} + 4\text{I}_2$
- 4)  $2\text{MnO}_4^- + 10\text{Cl}^- + 16\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} + 5\text{Cl}_2$
- 5)  $2\text{MnO}_4^- + 5\text{H}_2\text{O}_2 + 6\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} + 5\text{O}_2$
- 6)  $2\text{IO}_3^- + 10\text{I}^- + 12\text{H}^+ \rightarrow 6\text{I}_2 + 6\text{H}_2\text{O}$

### TASK 4 – Redox reactions or not?

Equation	Redox reaction?	Disproportionation on reaction?	Species oxidised	Species reduced	Oxidising agent	Reducing agent
$\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$	✓		CO	$\text{Fe}_2\text{O}_3$	$\text{Fe}_2\text{O}_3$	CO
$\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$	✓		Mg	HCl	HCl	Mg
$\text{MgO} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2\text{O}$	✗					
$\text{Al}_2\text{O}_3 + 2\text{Fe} \rightarrow 2\text{Al} + \text{Fe}_2\text{O}_3$	✓		Fe	$\text{Al}_2\text{O}_3$	$\text{Al}_2\text{O}_3$	Fe
$[\text{Co}(\text{H}_2\text{O})_6]^{2+} + 4\text{Cl}^- \rightarrow [\text{CoCl}_4]^{2-} + 6\text{H}_2\text{O}$	✗					
$\text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{NaOH}$	✗					
$2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$	✓	✓	$\text{H}_2\text{O}_2$	$\text{H}_2\text{O}_2$	$\text{H}_2\text{O}_2$	$\text{H}_2\text{O}_2$
$2\text{NaBr} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{HBr}$	✗					
$\text{Cl}_2 + 2\text{NaOH} \rightarrow \text{NaCl} + \text{NaOCl} + \text{H}_2\text{O}$	✓	✓	$\text{Cl}_2$	$\text{Cl}_2$	$\text{Cl}_2$	$\text{Cl}_2$

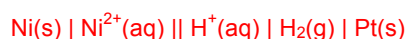
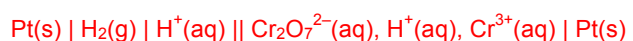
## TASK 5 – General AS Redox Questions

- 1) a) i)  $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$   
 ii)  $\text{N}_2 + 6\text{H}_2\text{O} \rightarrow 2\text{NO}_3^- + 10\text{e}^- + 12\text{H}^+$   
 iii)  $\text{V}^{3+} + \text{H}_2\text{O} \rightarrow \text{VO}^{2+} + \text{e}^- + 2\text{H}^+$   
 iv)  $\text{H}_2\text{SO}_4 + 6\text{e}^- + 6\text{H}^+ \rightarrow \text{S} + 4\text{H}_2\text{O}$   
 v)  $\text{NO}_3^- + 8\text{e}^- + 10\text{H}^+ \rightarrow \text{NH}_4^+ + 3\text{H}_2\text{O}$   
 vi)  $2\text{BrO}_3^- + 10\text{e}^- + 12\text{H}^+ \rightarrow \text{Br}_2 + 6\text{H}_2\text{O}$
- b)  $\text{H}_2\text{SO}_4 + 6\text{Br}^- + 6\text{H}^+ \rightarrow \text{S} + 4\text{H}_2\text{O} + 3\text{Br}_2$   
 oxidised =  $\text{Br}^-$ , reduced =  $\text{H}_2\text{SO}_4$ , oxidising agent =  $\text{H}_2\text{SO}_4$ , reducing agent =  $\text{Br}^-$
- c)  $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$   
 $\text{HNO}_3 + \text{H}^+ + \text{e}^- \rightarrow \text{NO}_2 + \text{H}_2\text{O}$   
 $2\text{HNO}_3 + 2\text{H}^+ + \text{Cu} \rightarrow 2\text{NO}_2 + 2\text{H}_2\text{O} + \text{Cu}^{2+}$
- 2) State whether the following three reactions are redox reactions or not. For those that are redox reactions, clearly indicate any changes in oxidation state.
- a)  $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$  redox Zn 0  $\rightarrow$  +2, H +1  $\rightarrow$  0
- b)  $\text{CuO} + 2\text{HCl} \rightarrow \text{CuCl}_2 + \text{H}_2\text{O}$  not redox
- c)  $\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + 2\text{H}_2\text{O}$  redox Mn +4  $\rightarrow$  +2, Cl -1  $\rightarrow$  0
- d)  $\text{Cl}_2 + \text{H}_2\text{O} \rightarrow \text{HCl} + \text{HOCl}$  redox Cl 0  $\rightarrow$  -1 and +1 (disproportionation)
- 3) a) oxidising agent = takes away electrons from another species  
 reducing agent = adds electrons to another species
- b) oxidised = Mg, reduced =  $\text{H}^+$ , oxidising agent =  $\text{H}^+$ , reducing agent = Mg
- c)
- |                                   |                   |
|-----------------------------------|-------------------|
| $\text{Na}_2\text{SO}_4$          | Na +1, S +6, O -2 |
| $\text{Na}_2\text{S}_2\text{O}_3$ | Na +1, S +2, O -2 |
| $\text{Na}_2\text{SO}_3$          | Na +1, S +4, O -2 |
| $\text{S}_8$                      | S 0               |
| $\text{KClO}_3$                   | K +1, Cl +5, O -2 |
| $\text{KClO}$                     | K +1, Cl +1, O -2 |
| $\text{KH}$                       | K +1, H -1        |
| $\text{Na}_2\text{O}_2$           | Na +1, O -1       |
| $\text{Na}_2\text{O}$             | Na +1, O -2       |
- 4) a) i)  $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$   
 ii)  $\text{SO}_4^{2-} + 2\text{e}^- + 2\text{H}^+ \rightarrow \text{SO}_3^{2-} + \text{H}_2\text{O}$   
 iii)  $\text{H}_2\text{O}_2 \rightarrow \text{O}_2 + 2\text{e}^- + 2\text{H}^+$   
 iv)  $2\text{IO}_3^- + 10\text{I}^- + 12\text{H}^+ \rightarrow 6\text{I}_2 + 6\text{H}_2\text{O}$   
 v)  $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$
- b)  $2\text{IO}_3^- + 10\text{I}^- + 12\text{H}^+ \rightarrow 6\text{I}_2 + 6\text{H}_2\text{O}$
- 5) a)  $2\text{MnO}_4^- + 10\text{Cl}^- + 16\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} + 5\text{Cl}_2$   
 b)  $6\text{MnO}_4^- + 10\text{Cr}^{3+} + 11\text{H}_2\text{O} \rightarrow 6\text{Mn}^{2+} + 5\text{Cr}_2\text{O}_7^{2-} + 22\text{H}^+$

## TASK 6 – The effect of changing conditions on electrode potential

- a) 298K, 100kPa H<sub>2</sub>(g), 1.0 mol dm<sup>-3</sup> H<sup>+</sup>(aq)
- b) by definition
- c) i) negative, equilibrium shifts left to side with less gas molecules to reduce pressure of H<sub>2</sub>  
ii) no effect, allows for faster rate of transfer of electrons but has no effect on potential  
iii) no effect, as concentration of H<sup>+</sup> = 1.00 mol dm<sup>-3</sup>  
iv) positive, equilibrium shifts right to lower concentration of H<sup>+</sup>  
v) negative, equilibrium shifts left in endothermic direction to reduce temperature

## TASK 7 – Writing conventional representations of cells



## TASK 8 – Electrode potentials

- 1)  $E^\circ = -2.71 \text{ V}$
- 2)  $\text{emf} = -0.44 - +0.22 = -0.66 \text{ V}$
- 3) a)  $\text{emf} = -0.13 - -0.76 = +0.63 \text{ V}$   
b)  $E^\circ$  (Pb<sup>2+</sup>/Pb) would become more negative as equilibrium shifts left to increases concentration of Pb<sup>2+</sup> ions; therefore emf would decrease and become less positive
- 4) a)  $\text{emf} = +0.15 - -0.25 = +0.40 \text{ V}$   
b)  $\text{emf} = +0.80 - +0.54 = +0.26 \text{ V}$   
c)  $\text{emf} = +1.07 - +1.36 = -0.29 \text{ V}$
- 5)  $1.02 = +1.36 - E_L^\circ$        $E_L^\circ = 1.36 - 1.02 = 0.34 \text{ V}$
- 6) a)  $2.00 = E_R^\circ - -2.38$        $E_R^\circ = 2.00 - 2.38 = -0.38 \text{ V}$   
b)  $0.54 = -2.38 - E_L^\circ$        $E_L^\circ = -2.38 - 0.54 = -2.92 \text{ V}$   
c)  $-3.19 = E_R^\circ - +0.27$        $E_R^\circ = -3.19 + 0.27 = -2.92 \text{ V}$
- 7) a)  $\text{Cr(s)} \mid \text{Cr}^{2+}(\text{aq}) \parallel \text{Zn}^{2+}(\text{aq}) \mid \text{Zn(s)}$      $\text{emf} = +0.15 \text{ V}$   
b)  $\text{Cu(s)} \mid \text{Cu}^{2+}(\text{aq}) \parallel \text{Fe}^{3+}(\text{aq}), \text{Fe}^{2+}(\text{aq}) \mid \text{Pt(s)}$      $\text{emf} = +0.43 \text{ V}$   
c)  $\text{Pt(s)} \mid \text{Cl}^-(\text{aq}) \mid \text{Cl}_2(\text{g}) \parallel \text{MnO}_4^-(\text{aq}), \text{H}^+(\text{aq}), \text{Mn}^{2+}(\text{aq}) \mid \text{Pt(s)}$      $\text{emf} = +0.15 \text{ V}$

## TASK 9 – Using the electrochemical series

- 1) a)  $\text{Ni}^{2+} + \text{Zn} \rightarrow \text{Ni} + \text{Zn}^{2+}$   
 b)  $E^\circ(\text{Ni}^{2+}/\text{Ni}) > E^\circ(\text{Zn}^{2+}/\text{Zn})$  and therefore  $\text{Ni}^{2+}$  gains electrons from Zn
- 2)  $2\text{Ag}^+ + \text{Cu} \rightarrow 2\text{Ag} + \text{Cu}^{2+}$   
 $E^\circ(\text{Ag}^+/\text{Ag}) > E^\circ(\text{Cu}^{2+}/\text{Cu})$  and therefore  $\text{Ag}^+$  gains electrons from Cu
- 3)  $\text{MnO}_4^-$  will oxidise  $\text{Cl}^-$  to form  $\text{Cl}_2$  as  $E^\circ(\text{MnO}_4^-/\text{Mn}^{2+}) > E^\circ(\text{Cl}_2/\text{Cl}^-)$  and therefore  $\text{MnO}_4^-$  gains electrons from  $\text{Cl}^-$   
 $\text{Cr}_2\text{O}_7^{2-}$  will **not** oxidise  $\text{Cl}^-$  to form  $\text{Cl}_2$  as  $E^\circ(\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}) < E^\circ(\text{Cl}_2/\text{Cl}^-)$  and therefore  $\text{Cr}_2\text{O}_7^{2-}$  does not gain electrons from  $\text{Cl}^-$   
 $\text{Fe}^{3+}$  will **not** oxidise  $\text{Cl}^-$  to form  $\text{Cl}_2$  as  $E^\circ(\text{Fe}^{3+}/\text{Fe}^{2+}) < E^\circ(\text{Cl}_2/\text{Cl}^-)$  and therefore  $\text{Fe}^{3+}$  does not gain electrons from  $\text{Cl}^-$
- 4) a)  $\text{Pt(s)} \mid \text{H}_2(\text{g}) \mid \text{H}^+(\text{aq}) \parallel \text{Br}_2(\text{l}) \mid \text{Br}^-(\text{aq}) \mid \text{Pt(s)}$   
 $\text{emf} = +1.09 \text{ V}$   
 anode =  $\text{H}^+/\text{H}_2$   
 $\text{Br}_2 + \text{H}_2 \rightarrow 2\text{Br}^- + 2\text{H}^+$   
 $E^\circ(\text{Br}_2/\text{Br}^-) > E^\circ(\text{H}^+/\text{H}_2)$  and therefore  $\text{Br}_2$  gains electrons from  $\text{H}_2$
- b)  $\text{Cu(s)} \mid \text{Cu}^{2+}(\text{aq}) \parallel \text{Fe}^{3+}(\text{aq}), \text{Fe}^{2+}(\text{aq}) \mid \text{Pt(s)}$   
 $\text{emf} = +0.43 \text{ V}$   
 anode =  $\text{Cu}^{2+}/\text{Cu}$   
 $2\text{Fe}^{3+} + \text{Cu} \rightarrow 2\text{Fe}^{2+} + \text{Cu}^{2+}$   
 $E^\circ(\text{Fe}^{3+}/\text{Fe}^{2+}) > E^\circ(\text{Cu}^{2+}/\text{Cu})$  and therefore  $\text{Fe}^{3+}$  gains electrons from Cu
- 5) a) the equilibrium  $\text{Cu}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Cu(s)}$   
 shifts left if concentration of  $\text{Cu}^{2+}$  is  $0.500 \text{ mol dm}^{-3}$  and so  $E^\circ$  becomes less than  $+0.34 \text{ V}$
- b) right hand electrode is anode
- c) from right hand side to left hand side
- 6) a) Yes:  $E^\circ(\text{H}^+/\text{H}_2) > E^\circ(\text{Fe}^{2+}/\text{Fe})$  and therefore  $\text{H}^+$  gains electrons from Fe  
 $2\text{H}^+ + \text{Fe} \rightarrow \text{H}_2 + \text{Fe}^{2+}$
- b) No:  $E^\circ(\text{H}^+/\text{H}_2) < E^\circ(\text{Cu}^{2+}/\text{Cu})$  and therefore  $\text{H}^+$  cannot gain electrons from Cu
- c) No:  $E^\circ(\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}) < E^\circ(\text{Cl}_2/\text{Cl}^-)$  and therefore  $\text{Cr}_2\text{O}_7^{2-}$  cannot gain electrons from  $\text{Cl}^-$
- d) Yes:  $E^\circ(\text{MnO}_4^-/\text{Mn}^{2+}) > E^\circ(\text{Cl}_2/\text{Cl}^-)$  and therefore  $\text{MnO}_4^-$  gains electrons from  $\text{Cl}^-$   
 $2\text{MnO}_4^- + 10\text{Cl}^- + 16\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} + 5\text{Cl}_2$
- e) Yes:  $E^\circ(\text{V}^{3+}/\text{V}^{2+}) > E^\circ(\text{Mg}^{2+}/\text{Mg})$  and therefore  $\text{V}^{3+}$  gains electrons from Mg  
 $\text{Mg} + 2\text{V}^{3+} \rightarrow \text{Mg}^{2+} + 2\text{V}^{2+}$

## TASK 10 – Commercial cells

- 1) a)  $\text{emf} = +1.10 \text{ V}$   
b)  $\text{Ag}_2\text{O} + 2\text{H}^+ + \text{Zn} \rightarrow 2\text{Ag} + \text{H}_2\text{O} + \text{Zn}^{2+}$   
c)  $\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}) \parallel \text{Ag}_2\text{O(s)} \mid \text{H}^+(\text{aq}) \mid \text{Ag(s)}$   
d) chemicals would be used up; goes flat when one or more of reactants runs out
- 2) a)  $\text{NiO(OH)} + \text{MH} \rightarrow \text{Ni(OH)}_2 + \text{M}$   
b)  $\text{Ni(OH)}_2 + \text{M} \rightarrow \text{NiO(OH)} + \text{MH}$   
c) can reuse rather than throw away
- 3) a)  $\text{CH}_3\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 4\text{H}_2\text{O}$   
b)  $+1.21 \text{ V}$   
c) does not go flat; does not need re-charging