ELECTROCHEMICAL CELLS

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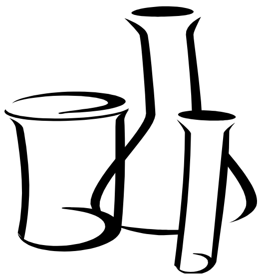

TASK 1 – Oxidation states

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

a) *Fe* in Fe 0 $FeCl_3$ +3 $FeCl_2$ +2 K_2FeO_4 +6 $[Fe(H_2O)_6]^{2+}$ +2

b) *Cl* in Cl_2 0 ClO_3^- +5 ClO^- +1 Cl_2O_7 +7   NaCl -1

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

SO_2	S +4	O -2
S	S 0	
SO_3	S +6	O -2
H_2S	H +1	S -2
NH_3	N -3	H +1
NO_2	N +4	O -2
N_2	N 0	
NO_3^-	N +5	O -2
Cl^-	Cl -1	
SO_4^{2-}	S +6	O -2

H_2O_2	H +1	O -1	
NaH	Na +1	H -1	
CO_3^{2-}	C +4	O -2	
Cr_2O_3	Cr +3	O -2	
CrO_3	Cr +6	O -2	
$KMnO_4$	K +1	Mn +7	O -2
K_2MnO_4	K +1	Mn +6	O -2
Cu_2O	Cu +1	O -2	
CuO	Cu +2	O -2	
$NaCuCl_2$	Na +1	Cu +1	Cl -1

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

a) $[Co(H_2O)_6]^{2+}$ Co +2

b) $[CrCl_6]^{3-}$ Cr +3

c) $[Co(NH_3)_6]Cl_2$ Co +2

d) $[Co(NH_3)_5Cl]Cl_2$ Co +3

e) $[FeO_4]^{2-}$ Fe +6

f) $[Mn(CN)_6]^{4-}$ Mn +2

g) $[Ni(CO)_4]$ Ni 0

h) $[Ni(EDTA)]^{2-}$ Ni +2

i) $K_2[CoCl_4]$ Co +2

j) $K_3[AuF_6]$ Au +3

k) $(NH_4)_2[IrCl_6]$ Ir +4

l) $Na[Mn(CO)_5]$ Mn -1

TASK 2 – Writing half equations

a) $2I^- \rightarrow I_2 + 2e^-$

b) $Fe^{2+} \rightarrow Fe^{3+} + e^-$

c) $Fe \rightarrow Fe^{2+} + 2e^-$

d) $2Cl^- \rightarrow Cl_2 + 2e^-$

e) $SO_4^{2-} + 2e^- + 4H^+ \rightarrow SO_2 + 2H_2O$

f) $VO^{2+} + H_2O \rightarrow VO_2^+ + e^- + 2H^+$

g) $SO_4^{2-} + 8e^- + 10H^+ \rightarrow H_2S + 4H_2O$

h) $H_2O_2 \rightarrow O_2 + 2e^- + 2H^+$

i) $Cr_2O_7^{2-} + 6e^- + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$

j) $C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$

k) $Hg_2^{2+} + 2e^- \rightarrow 2Hg$

l) $2IO_3^- + 10e^- + 12H^+ \rightarrow I_2 + 6H_2O$

TASK 3 – Combining half equations

1) $Zn + 2Fe^{3+} \rightarrow Zn^{2+} + 2Fe^{2+}$
2) $Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O + 6Fe^{3+}$
3) $H_2SO_4 + 8I^- + 8H^+ \rightarrow H_2S + 4H_2O + 4I_2$
4) $2MnO_4^- + 10Cl^- + 16H^+ \rightarrow 2Mn^{2+} + 8H_2O + 5Cl_2$
5) $2MnO_4^- + 5H_2O_2 + 6H^+ \rightarrow 2Mn^{2+} + 8H_2O + 5O_2$
6) $2IO_3^- + 10I^- + 12H^+ \rightarrow 6I_2 + 6H_2O$

TASK 4 – Redox reactions or not?

Equation	Redox reaction?	Disproportionation reaction?	Species oxidised	Species reduced	Oxidising agent	Reducing agent
$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$	✓		CO	Fe_2O_3	Fe_2O_3	CO
$Mg + 2HCl \rightarrow MgCl_2 + H_2$	✓		Mg	HCl	HCl	Mg
$MgO + 2HCl \rightarrow MgCl_2 + H_2O$	✗					
$Al_2O_3 + 2Fe \rightarrow 2Al + Fe_2O_3$	✓		Fe	Al_2O_3	Al_2O_3	Fe
$[Co(H_2O)_6]^{2+} + 4Cl^- \rightarrow [CoCl_4]^{2-} + 6H_2O$	✗					
$Na_2O + H_2O \rightarrow 2NaOH$	✗					
$2H_2O_2 \rightarrow 2H_2O + O_2$	✓	✓	H_2O_2	H_2O_2	H_2O_2	H_2O_2
$2NaBr + H_2SO_4 \rightarrow Na_2SO_4 + 2HBr$	✗					
$Cl_2 + 2NaOH \rightarrow NaCl + NaOCl + H_2O$	✓	✓	Cl_2	Cl_2	Cl_2	Cl_2

TASK 5 – General AS Redox Questions

1) a) i) $2Br^- \rightarrow Br_2 + 2e^-$
ii) $N_2 + 6H_2O \rightarrow 2NO_3^- + 10e^- + 12H^+$
iii) $V^{3+} + H_2O \rightarrow VO^{2+} + e^- + 2H^+$
iv) $H_2SO_4 + 6e^- + 6H^+ \rightarrow S + 4H_2O$
v) $NO_3^- + 8e^- + 10H^+ \rightarrow NH_4^+ + 3H_2O$
vi) $2BrO_3^- + 10e^- + 12H^+ \rightarrow Br_2 + 6H_2O$

b) $H_2SO_4 + 6Br^- + 6H^+ \rightarrow S + 4H_2O + 3Br_2$
oxidised = Br^-, reduced = H_2SO_4, oxidising agent = H_2SO_4, reducing agent = Br^-

c) $Cu \rightarrow Cu^{2+} + 2e^-$
$HNO_3 + H^+ + e^- \rightarrow NO_2 + H_2O$
$2HNO_3 + 2H^+ + Cu \rightarrow 2NO_2 + 2H_2O + Cu^{2+}$

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) $Zn + 2HCl \rightarrow ZnCl_2 + H_2$ — redox Zn 0 → +2, H +1 → 0

b) $CuO + 2HCl \rightarrow CuCl_2 + H_2O$ — not redox

c) $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$ — redox Mn +4 → +2, Cl -1 → 0

d) $Cl_2 + H_2O \rightarrow HCl + HOCl$ — redox Cl 0 → -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 = H^+, oxidising agent = H^+, reducing agent = Mg

c)

Na_2SO_4	Na +1, S +6, O -2
$Na_2S_2O_3$	Na +1, S +2, O -2
Na_2SO_3	Na +1, S +4, O -2
S_8	S 0
$KClO_3$	K +1, Cl +5, O -2
$KClO$	K +1, Cl +1, O -2
KH	K +1, H -1
Na_2O_2	Na +1, O -1
Na_2O	Na +1, O -2

4) a) i) $2H^+ + 2e^- \rightarrow H_2$
ii) $SO_4^{2-} + 2e^- + 2H^+ \rightarrow SO_3^{2-} + H_2O$
iii) $H_2O_2 \rightarrow O_2 + 2e^- + 2H^+$
iv) $2IO_3^- + 10I^- + 12H^+ \rightarrow 6I_2 + 6H_2O$
v) $2I^- \rightarrow I_2 + 2e^-$

b) $2IO_3^- + 10I^- + 12H^+ \rightarrow 6I_2 + 6H_2O$

5) a) $2MnO_4^- + 10Cl^- + 16H^+ \rightarrow 2Mn^{2+} + 8H_2O + 5Cl_2$
b) $6MnO_4^- + 10Cr^{3+} + 11H_2O \rightarrow 6Mn^{2+} + 5Cr_2O_7^{2-} + 22H^+$

TASK 6 – The effect of changing conditions on electrode potential

a) 298K, 100kPa $H_2(g)$, 1.0 mol dm^{-3} $H^+(aq)$

b) by definition

c) i) negative, equilibrium shifts left to side with less gas molecules to reduce pressure of H_2

ii) no effect, allows for faster rate of transfer of electrons but has no effect on potential

iii) no effect, as concentration of H^+ = 1.00 mol dm^{-3}

iv) positive, equilibrium shifts right to lower concentration of H^+

v) negative, equilibrium shifts left in endothermic direction to reduce temperature

TASK 7 – Writing conventional representations of cells

$Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$

$Pt(s) \mid H_2(g) \mid H^+(aq) \parallel Cr_2O_7^{2-}(aq), H^+(aq), Cr^{3+}(aq) \mid Pt(s)$

$Ni(s) \mid Ni^{2+}(aq) \parallel H^+(aq) \mid H_2(g) \mid Pt(s)$

$Al(s) \mid Al^{3+}(aq) \parallel Pb^{2+}(aq) \mid Pb(s)$

$Fe(s) \mid Fe^{2+}(aq) \parallel MnO_4^-(aq), H^+(aq), Mn^{2+}(aq) \mid Pt(s)$

$Ag(s) \mid Ag^+(aq) \parallel Zn^{2+}(aq) \mid Zn(s)$

TASK 8 – Electrode potentials

1) $E^o = -2.71$ V

2) emf = $-0.44 - +0.22 = -0.66$ V

3) a) emf = $-0.13 - -0.76 = +0.63$ V

b) E^o (Pb^{2+}/Pb) would become more negative as equilibrium shifts left to increases concentration of Pb^{2+} ions; therefore emf would decrease and become less positive

4) a) emf = $+0.15 - -0.25 = +0.40$ V

b) emf = $+0.80 - +0.54 = +0.26$ V

c) emf = $+1.07 - +1.36 = -0.29$ V

5) $1.02 = +1.36 - E_L^o$ $\quad E_L^o = 1.36 - 1.02 = 0.34$ V

6) a) $2.00 = E_R^o - -2.38$ $\quad E_R^o = 2.00 - 2.38 = -0.38$ V

b) $0.54 = -2.38 - E_L^o$ $\quad E_L^o = -2.38 - 0.54 = -2.92$ V

c) $-3.19 = E_R^o - +0.27$ $\quad E_R^o = -3.19 + 0.27 = -2.92$ V

7) a) $Cr(s) \mid Cr^{2+}(aq) \parallel Zn^{2+}(aq) \mid Zn(s)$ emf = +0.15 V

b) $Cu(s) \mid Cu^{2+}(aq) \parallel Fe^{3+}(aq), Fe^{2+}(aq) \mid Pt(s)$ emf = +0.43 V

c) $Pt(s) \mid Cl^-(aq) \mid Cl_2(g) \parallel MnO_4^-(aq), H^+(aq), Mn^{2+}(aq) \mid Pt(s)$ emf = +0.15 V

TASK 9 – Using the electrochemical series

1) a) $Ni^{2+} + Zn \rightarrow Ni + Zn^{2+}$

b) $E°(Ni^{2+}/Ni) > E°(Zn^{2+}/Zn)$ and therefore Ni^{2+} gains electrons from Zn

2) $2Ag^{+} + Cu \rightarrow 2Ag + Cu^{2+}$

$E°(Ag^{+}/Ag) > E°(Cu^{2+}/Cu)$ and therefore Ag^{+} gains electrons from Cu

3) MnO_4^{-} will oxidise Cl^{-} to form Cl_2 as $E°(MnO_4^{-}/Mn^{2+}) > E°(Cl_2/ Cl^{-})$ and therefore MnO_4^{-} gains electrons from Cl^{-}

$Cr_2O_7^{2-}$ will **not** oxidise Cl^{-} to form Cl_2 as $E°(Cr_2O_7^{2-}/Cr^{3+}) < E°(Cl_2/ Cl^{-})$ and therefore $Cr_2O_7^{2-}$ does not gain electrons from Cl^{-}

Fe^{3+} will **not** oxidise Cl^{-} to form Cl_2 as $E°(Fe^{3+}/Fe^{2+}) < E°(Cl_2/ Cl^{-})$ and therefore Fe^{3+} does not gain electrons from Cl^{-}

4) a) Pt(s) | H_2(g) | H^{+}(aq) || Br_2(l) | Br^{-}(aq) | Pt(s)

emf = +1.09 V

anode = H^{+}/H_2

$Br_2 + H_2 \rightarrow 2Br^{-} + 2H^{+}$

$E°(Br_2/Br^{-}) > E°(H^{+}/H_2)$ and therefore Br_2 gains electrons from H_2

b) Cu(s) | Cu^{2+}(aq) || Fe^{3+}(aq), Fe^{2+}(aq) | Pt(s)

emf = +0.43 V

anode = Cu^{2+}/Cu

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

$E°(Fe^{3+}/Fe^{2+}) > E°(Cu^{2+}/Cu)$ and therefore Fe^{3+} gains electrons from Cu

5) a) the equilibrium $Cu^{2+}(aq) + 2e^{-} \rightleftharpoons Cu(s)$

shifts left if concentration of Cu^{2+} is 0.500 mol dm^{-3} and so E° becomes less than +0.34 V

b) right hand electrode is anode

c) from right hand side to left hand side

6) a) Yes: $E°(H^{+}/H_2) > E°(Fe^{2+}/Fe)$ and therefore H^{+} gains electrons from Fe

$2H^{+} + Fe \rightarrow H_2 + Fe^{2+}$

b) No: $E°(H^{+}/H_2) < E°(Cu^{2+}/Cu)$ and therefore H^{+} cannot gain electrons from Cu

c) No: $E°(Cr_2O_7^{2-}/Cr^{3+}) < E°(Cl_2/ Cl^{-})$ and therefore $Cr_2O_7^{2-}$ cannot gain electrons from Cl^{-}

d) Yes: $E°(MnO_4^{-}/Mn^{2+}) > E°(Cl_2/ Cl^{-})$ and therefore MnO_4^{-} gains electrons from Cl^{-}

$2MnO_4^{-} + 10Cl^{-} + 16H^{+} \rightarrow 2Mn^{2+} + 8H_2O + 5Cl_2$

e) Yes: $E°(V^{3+}/V^{2+}) > E°(Mg^{2+}/Mg)$ and therefore V^{3+} gains electrons from Mg

$Mg + 2V^{3+} \rightarrow Mg^{2+} + 2V^{2+}$

TASK 10 – Commercial cells

1) a) emf = +1.10 V
   b) $Ag_2O + 2H^+ + Zn \longrightarrow 2Ag + H_2O + Zn^{2+}$
   c) $Zn(s) \mid Zn^{2+}(aq) \parallel Ag_2O(s) \mid H^+(aq) \mid Ag(s)$
   d) chemicals would be used up; goes flat when one or more of reactants runs out

2) a) $NiO(OH) + MH \longrightarrow Ni(OH)_2 + M$
   b) $Ni(OH)_2 + M \longrightarrow NiO(OH) + MH$
   c) can reuse rather than throw away

3) a) $CH_3OH + 1\frac{1}{2}O_2 \longrightarrow CO_2 + 2H_2O$
   b) +1.21 V
   c) does not go flat; does not need re-charging