1) What is the oxidation number of the underlined element?

NO $\quad$ **N**O_3^- $\quad$ **Cl**O_3^- $\quad$ **V**O^{2+}

+2 $\quad$ +5 $\quad$ +5 $\quad$ +4

2) Which of the following elements is getting oxidized and which is getting reduced? Also indicate the elements is an oxidizing and which is a reducing agents the above reactions?

a) $4HCl + O_2 \longrightarrow 2Cl_2 + 2H_2O$

−1 $\quad$ 0 $\quad$ 0 $\quad$ −2

oxidized. $\quad$ reduced

reducing agent $\quad$ oxidizing agent.

b) $CH_4 + 4NO \longrightarrow 2N_2 + CO_2 + 2H_2O$

−4 $\quad$ +2 $\quad$ 0 $\quad$ +4 $\quad$ +1

oxidized. $\quad$ reduction

reducing agent. $\quad$ oxidizing agent

3) Show the step by step process of balancing the redox reactions below in an acidic medium.

a) $Fe^{3+}(aq) + NH_2OH_2^+(aq) \longrightarrow Fe^{2+}(aq) + N_2O(g)$

$(Fe^{3+} \longrightarrow Fe^{2+} + 1e^-) \times 4$ $\quad$ $2NH_2OH_2^{\oplus} \longrightarrow N_2O + H_2O + 6H^{\oplus} + 4e^-$

$$2NH_2OH_2^{\oplus} + 4Fe^{3+} \longrightarrow 4Fe^{2+} + N_2O + H_2O + 6H^{\oplus}$$

b) $Fe(OH)_2(s) + O_2(g) \longrightarrow Fe(OH)_3(s)$

$(Fe(OH)_2 + H_2O \longrightarrow Fe(OH)_3 + H^{\oplus} + 1e^-) \times 4.$ $\quad$ $O_2 \longrightarrow 2H_2O$, $4H^{\oplus}$, $+4e^-$

$4Fe(OH)_2 + 4H_2O \longrightarrow 4Fe(OH)_3 + 4H^{\oplus} + 4e^-$

$O_2 + 4H^{\oplus} + 4e^- \longrightarrow 2H_2O$

$$4Fe(OH)_2 + 2H_2O + O_2 + 4H^{\oplus} \longrightarrow 4Fe(OH)_3$$

4) Show the step by step process of balancing the redox reactions below in a basic medium.

a) $Ag(s) + NO_3^-(aq) \longrightarrow Ag^+(aq) + NO(g)$

$(Ag \rightarrow Ag^{\oplus} + 1e^-) \times 3 \quad NO_3^{\ominus} \rightarrow NO + 2H_2O$

$+4H^{\oplus}$
$+3e^-$

$3Ag + NO_3^{\ominus} + 4H^{\oplus} \rightarrow 3Ag^{\oplus} + NO + 2H_2O$

$+4OH^-$ $\qquad +4OH^-$

$3Ag + NO_3^{\ominus} + 4H_2O \rightarrow 3Ag^{\oplus} + NO + 2H_2O$

$3Ag + NO_3^{\ominus} + 2H_2O \rightarrow 3Ag^{\oplus} + NO$

b) $S_8(s) \longrightarrow S_2O_3^{2-}(aq) + S^{2-}(aq)$

$S_8 \rightarrow 4S_2O_3^{2-} \qquad S_8 \rightarrow 8S^{2-}$

$+12H_2O \qquad +16e^-$

$+24H^{\oplus}$
$+16e^-$

$2S_8 + 12H_2O \rightarrow 4S_2O_3^{2-} + 8S^{2-} + 24H^{\oplus}$

$+24OH^{\ominus}$

$+24OH^{\ominus}$

$24H_2O$

$2S_8 + 24OH^{\ominus} \rightarrow 4S_2O_3^{2-} + 8S^{2-} + 12H_2O$

5) Below is the voltaic cell notation of a galvanic cell. Write the half reactions and which is the cathode and anode in the cell.

$Cu(s) \mid Cu^{2+}(0.10\ M) \parallel Ag_2CrO_4\ (satd\ aq) \mid Ag(s)$

(A) $Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^-$

(C) $Ag_2CrO_4 \longrightarrow Ag$.

$\hookrightarrow Ag^{\oplus} + 1e^- \longrightarrow Ag$.

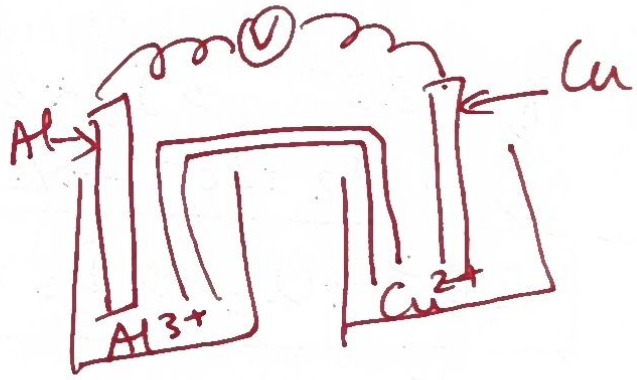
6) Write the half reactions for the redox reaction given below. Write the voltaic cell notation and sketch a galvanic cell for the reaction.

$2Al\ (s) + 3\ Cu^{2+}(aq) \longrightarrow 2\ Al^{3+}\ (aq) + 3\ Cu(s)$

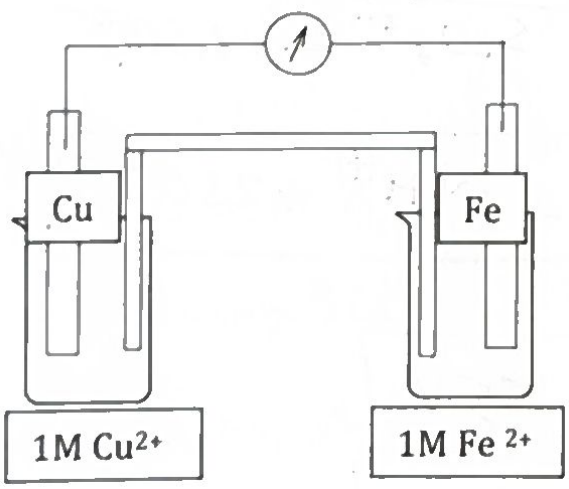
(A) $Al \longrightarrow Al^{3+} + 3e^-$

(C) $Cu^{2+} + 2e^- \longrightarrow Cu$

$Al(s) \mid Al^{3+}_{(aq)}(1.0M) \parallel Cu^{2+}(1.0M) \mid Cu_{(s)}$



7) Below is the diagram of a galvanic cell. Write the half equations and redox equation associated with this cell.



$Cu \longrightarrow Cu^{2+} + 2e^-$

$Fe^{2+} + 2e^- \longrightarrow Fe$

8) The E^{o}_{cell} for the following reaction is 1.334V.

$3V(s) + 2SbO^{+}(aq) + 4H^{+}(aq) \longrightarrow 3V^{2+}(aq) + 2Sb(s) + 2H_2O(l)$

anode →

If reduction potential, E^{o}, $V^{2+}_{(aq)}/V_{(s)}$ = -1.13V, then what is the E^{o} for the reaction below?

$SbO^{+}(aq) + 2H^{+}(aq) + 3e^{-} \longrightarrow Sb(s) + H_2O(l)$ E^{o} = ? (Ans: 0.20 V)

$E_{cell} = E_{cath} - E_{anode}$.

$1.334 = cath - (-1.13)$

$cath. = \boxed{0.20V}$.

9) Write the equations for the half reactions and overall cell reaction and calculate the E^{o}_{cell} for each of the voltaic cells shown below. (*Hint: Use the redox potential from the power point data tables.*)

a) $I^{-}(aq) \mid I_2(s) \,||\, Cl_2(g) \mid Cl^{-}(aq) \mid Pt$ (Ans: 0.82 V)

(A) $2I^{\ominus} \rightarrow I_2 + 2e$ 0.535V

(C) $Cl_2 + 2e^{-} \rightarrow 2Cl^{\ominus}$ 1.358V

$\boxed{2I^{\ominus} + Cl_2 \rightarrow I_2 + 2Cl^{\ominus}}$

$E_{cell} = cath - anode = 1.358 - 0.535 = \boxed{0.823V}$.

b) $Pb^{2+}(aq), H^{+}(aq) \mid PbO_2(s) \,||\, S_2O_8^{2-}(aq), SO_4^{2-}(aq) \mid Pt$ (Ans: 0.32 V)

inverse.

$PbO_2 + 4H^{\oplus} + \cancel{2e^{-}} \rightarrow Pb^{2+} + 2H_2O$ anode 1.455V

$S_2O_8^{2-} \rightarrow 2SO_4^{2-}$ cathode 2.01 V

$\cancel{+2e^{-}}$

$S_2O_8^{2-} \rightarrow 2SO_4^{2-}$

$Pb^{2+} + 2H_2O \rightarrow PbO_2 + 4H^{+}$

$\boxed{Pb^{2+} + 2H_2O + S_2O_8^{2-} - PbO_2 + 4H^{\oplus} + 2SO_4^{2-}}$

E_{cell} = cath - anode.

$2.01 - 1.455 = \boxed{0.55V}$.

10) Write the half reactions for the following redox equations and show it in a voltaic cell representation and calculate the E^{o}_{cell}.

a) $Fe^{3+}(aq) + Sn^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Sn^{4+}(aq)$ (Ans: 0.62 V)

$Sn^{2+} \longrightarrow Sn^{4+} + 2e$ (A) 0.154 V

$Fe^{3+} + 1e^- \longrightarrow Fe^{2+}$ (C) 0.771 V

$Sn^{2+}_{(aq)}(1M) \mid Sn^{4+}_{(aq)}(1M) \,\|\, Fe^{3+}_{(aq)}(1M) \mid Fe^{2+}_{(aq)}(1M)$.

E_{cell} = cath − anode = 0.771 − 0.154 = **0.617V**

b) $Cu(s) + H^+(aq) + NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + H_2O(l) + NO(g)$ (Ans: 0.62 V)

$Cu \longrightarrow Cu^{2+} + 2e^-$ 0.340V

$4H^+ + NO_3^- + 3e^- \longrightarrow NO + 2H_2O$ · 0.956V

$Cu_{(s)} \mid Cu^{2+}_{(aq)}(1M) \,\|\, NO_{3\,(aq)}^{\ominus}(1M) \mid NO_{(g)}(1atm) \mid Pt_{(s)}$

E_{cell} = cath − anode = 0.956 − 0.340 = **0.616V**

11) Predict whether the following reactions will occur as written or no, i.e. if they are spontaneous. (*Hint: Use the redox potential from the power point data tables to find if the E_{cell} is positive*)

a) $Sn^{4+}(aq) + 2I^-(aq) \longrightarrow Sn^{2+}(aq) + I_2(s)$ (Ans: -0.39 V, non-spont.)

$Sn^{4+} \longrightarrow Sn^{2+}$ (C) 0.154V

$2I^{\ominus} \longrightarrow I_2$ (A) 0.535V

E_{cell} = cath − anode

0.15V − 0.535 = **−.385V Non spont**.

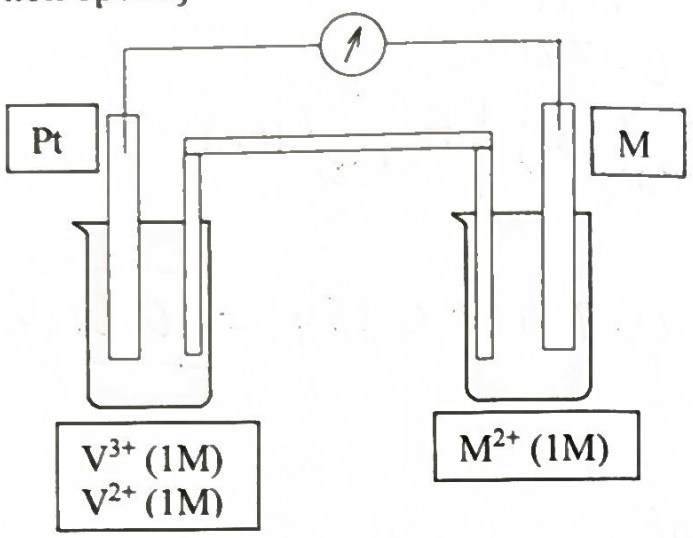
* b) $2MnO_2(s) + 3ClO^-(aq) + 2OH^-(aq) \longrightarrow 2MnO_4^-(aq) + 3Cl^-(aq) + H_2O(l)$ (Ans: 0.31 V, spont.)

$MnO_2 \longrightarrow MnO_4^-$ (A) 0.60 V

$ClO^{\ominus} \longrightarrow 3Cl^{\ominus}$ (C) 0.89V

E_{cell} = cath − anod = 0.89 − 0.60 = **0.29V spont**.

12) For the voltaic cell shown below, write an equation for the cell reaction that occurs and determine the voltage if the metal, M, is a) Sn and b) Zn. Are the reactions spontaneous? (*Use the redox potentials from the power point data tables*). (Ans: a) 0.118 V, spont.; b) -0.508 V, non-spont.)



anode needs to be more -ve for E_{cell} to be spont.

(a) $V^{2+} \longrightarrow V^{3+}$ (A) -0.255

$Sn^{2+} \longrightarrow Sn$ (C) -0.137

E_{cell} = cath − anode

= $-0.137-(-0.255)$

= **0.118 V spont.**

(b) $V^{2+} \longrightarrow V^{3+}$ (A) -0.255

$Zn^{2+} \longrightarrow Zn$ (C) -0.763

E_{cell} = cath − anode.

= $-0.763-(-0.255)$

= **−0.508 V non spont**

13) Are the following reactions possible – yes or no? Explain your answer. (*Hint: Use the redox potential from the power point data tables to find if the E_{cell} is positive*)

a) Displacement of $Cd^{2+}(aq)$ by Al(s). (Ans: 1.273 V, spont.)

$Al \rightarrow Al^{3+}$ (A) -1.676

$Cd^{2+} \rightarrow Cd$ (C) -0.403

Spont.

E_{cell} = cath − anode

= $-0.403-(-1.676)$

= **1.273 V spont**

b) Oxidation of $Cl^-(aq)$ to Cl_2 by $Br_2(l)$. (Ans: -0.293 V, non-spont.)

$Cl^{\ominus} \longrightarrow Cl_2$ (A) 1.358

$Br_2 \rightarrow Br^{\ominus}$ (C) 1.065

non spont.

cathode needs to be more +ve for E_{cell} to be spont.

E_{cell} = cath − anode.

= $1.065-1.358$

= **−0.293 V non spont**