Electrochemistry – Practice **Name:** Key

1) What is the oxidation number of the underlined element?
   <u>N</u>O   <u>N</u>O_3^- <u>Cl</u>O_3^- <u>V</u>O^{2+}

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

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

3) Balance the following redox reactions in an acidic medium.
   a) $Fe^{3+}(aq) + NH_2OH_2^+(aq) \longrightarrow Fe^{2+}(aq) + N_2O(g)$

$(\cancel{e^-} + Fe^{3+} \longrightarrow Fe^{2+}) \times 4$

$2NH_2OH_2^{+} \longrightarrow N_2O + H_2O + 6H^+ + \cancel{4e^-}$

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

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

$(Ag \longrightarrow Ag^{+} + \cancel{e^-}) \times 3$

$NO_3^{-} + 4H^+ + \cancel{3e^-} \longrightarrow NO + 2H_2O$

$$\boxed{3Ag + NO_3^{-} + 4H^{+} \longrightarrow 3Ag^{+} + NO + 2H_2O}$$

4) Balance the following redox reactions in a basic medium.

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

$\left(\begin{array}{l} Fe(OH)_2 \longrightarrow Fe(OH)_3 \\ + H_2O \qquad\qquad + H^{\oplus} \\ \qquad\qquad\qquad + e^- \end{array} \right) \times 4$. $\quad \begin{array}{l} O_2 \longrightarrow 2H_2O \\ 4H^{\oplus} \\ + 4e^- \end{array}$

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

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

$4\,Fe(OH)_2 + 4H_2O \longrightarrow 4\,Fe(OH)_3 + 2H_2O$

no $H^{\oplus}$ leftover.

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

$\begin{array}{l} S_8 \longrightarrow 4S_2O_3^{2-} \\ + 12H_2O \qquad + 24H^{\oplus} \\ \qquad\qquad\quad + 16e^- \end{array} \qquad \begin{array}{l} S_8 \longrightarrow 8S^{2-} \\ + 16e^- \end{array}$

$S_8 + 12H_2O \longrightarrow 4S_2O_3^{2-} + 24H^{\oplus} + \cancel{16e^-}$

$\cancel{16e^-} + S_8 \longrightarrow 8S^{2-}$

$2S_8 + 12H_2O + 24OH^{\ominus} \longrightarrow 4S_2O_3^{2-} + 8S^{2-} + 24H^{\oplus} + 24OH^{\ominus} \Big\}\; 24H_2O$

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

5) Calculate the E^o for the reactions as indicated.

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

anode

$E^o_{cell} = 1.334V$

$SbO^+(aq) \;+\; 2H^+(aq) \;+ 3e^- \longrightarrow Sb(s) \;+\; H_2O(l) \qquad E^o_{cell} = ?$

cath.

If $E^o\, V^{2+}_{(aq)}/V_{(s)} = -1.13V$

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

$1.334 = E_{cath} - (-1.13)$

$E_{cath} = 1.334 - 1.13 = \boxed{0.20V}$

6) Write the equation for the half reactions and overall cell reaction and calculate the E^o_{cell} for each of the voltaic cells shown below.

a) $Pt \mid I_2(s) \mid I^-(aq) \parallel Cl^-(aq) \mid Cl_2(g) \mid Pt$

$2I^{\ominus} \rightarrow I_2 + 2e^-$ anode = 0.535

$Cl_2 + 2e^- \rightarrow 2Cl^{\ominus}$ cathode = 1.358

$Cl_2 + \cancel{2e^-} \rightarrow 2Cl^{\ominus}$

$2I^{\ominus} \rightarrow I_2 + \cancel{2e^-}$

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

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

b) $Pt \mid PbO_2(s) \mid Pb^{2+}(aq), H^+(aq) \parallel S_2O_8^{2-}(aq), SO_4^{2-}(aq) \mid Pt$

$PbO_2 + 4H^{\oplus} + 2e^- \rightarrow Pb^{2+} + 2H_2O$ (inverse)

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

$S_2O_8^{2-} + 2e^- \rightarrow 2SO_4^{2-}$ cath. 2.01

$S_2O_8^{2-} + \cancel{2e^-} \rightarrow 2SO_4^{2-}$

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

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

E_{cell} = 2.01 − 1.455
= $\boxed{0.55V}$

*7) Write the half reactions for the following redox equations and show it in a voltaic cell representation. and calc. the Ecell.

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

$(Fe^{3+} + 1e^- \rightarrow Fe^{2+} + \cancel{1e^-}) \times 2$

$Sn^{2+} + \cancel{2e^-} \rightarrow Sn^{4+} + 2e^-$

cath = 0.771
anode = 0.154

$Pt \mid Sn^{2+}_{(aq)} \mid Sn^{4+}_{(aq)} \parallel Fe^{2+}_{(aq)} \mid Fe^{3+}_{(aq)} \mid Pt$

E_{cell} = 0.771 − 0.154
= $\boxed{0.617V}$

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

$(Cu \rightarrow Cu^{2+} + \cancel{2e}) \times 3$

$(4H^{\oplus} + NO_3^{\ominus} + \cancel{3e^-} \rightarrow 2H_2O + NO) \times 2$

anode = 0.340V
cath : 0.956

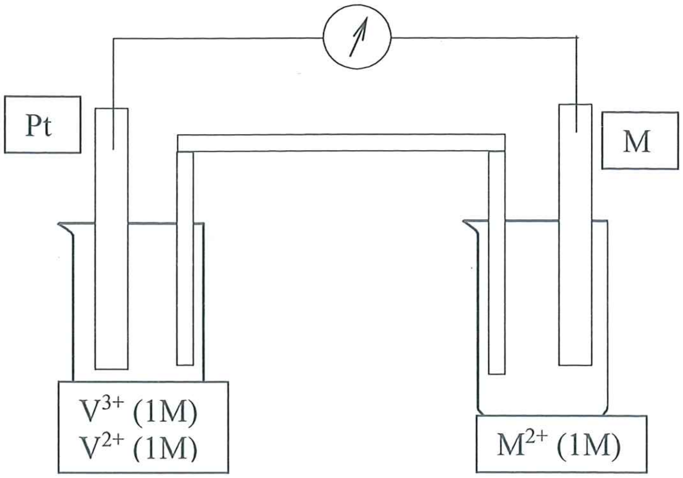
$3Cu + 8H^{\oplus} + 2NO_3^{\ominus} \rightarrow 3Cu^{2+} + 4H_2O + 2NO$.

$Cu_{(s)} \mid Cu^{2+}_{(aq)} \parallel H^{\oplus}_{(aq)} \mid NO_{3(aq)}^{\ominus} \mid NO_{(g)} \mid Pt$

$NO \mid NO_3^{\ominus}$

E_{cell} = 0.956 − 0.340
= $\boxed{0.616V}$

8) 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. Is the rxn spontaneous?



(a) $V^{2+} \rightarrow V^{3+} + e^-$ anode -0.255

$Sn^{2+} + 2e^- \rightarrow Sn$ cathode -0.137

E_{cell} = cath − anode

$= -0.137 - (-0.255) = \boxed{0.118V}$

SPONTANEOUS.

(b) $V^{2+} \rightarrow V^{3+} + e^-$ anode $= -0.255$

$Zn + 2e^- \rightarrow Zn^{2+}$ cath $= -0.763$.

E_{cell} = cath − anode

$= -0.763 - (-0.255)$

$= \boxed{-0.508V}$ $\boxed{\text{NON SPONT.}}$

9) Predict whether the following reactions will occur as written or no, i.e. if they are spontaneous.

a) $Sn^{4+}(aq) + 2I^-(aq) \longrightarrow Sn^{2+}(aq) + I_2(s)$

cath $= 0.154$ anode $= 0.535$

E_{cell} = cath − anode $= 0.154 - 0.535 = \boxed{-0.381V \text{ NON SPONT.}}$

b) $2MnO_2(s) + 3ClO^-(aq) + 2OH^-(aq) \longrightarrow 2MnO^{4-}(aq) + 3Cl^-(aq) + H_2O(l)$

cathode $= 0.890$

anode $= 0.60$

E_{cell} = cath − anode $= 0.89 - 0.60 = \boxed{0.29V \text{ SPONT.}}$

10) Are the following reactions possible – yes or no?

a) Displacement of Cd^{2+}(aq) by Al(s)

$Al \rightarrow Al^{3+}$ anode. -1.676

$Cd^{2+} \rightarrow Cd$ cathode -0.403

$E_{cell} = -0.403 - (-1.676)$

$= \boxed{1.273V \text{ SPONT.}}$

b) Oxidation of Cl^-(aq) to Cl_2 by Br_2(l)

$Cl^- \rightarrow Cl_2$ anode $= 1.358$

$Br_2 \rightarrow Br^-$ cath $= 1.065$

$E_{cell} = 1.065 - 1.358$

$= \boxed{-0.293V \text{ NON. SPONT}}$