REVIEW QUESTIONS
Chapter 19

1. For each of the following unbalanced equations, (i) write the half-reactions for oxidation and reduction, and (ii) balance the overall equation in acidic solution using the half-reaction method.

a) Oxidation numbers: +4, –1, +2, 0

$$MnO_2 + Cl^- \rightarrow Mn^{2+} + Cl_2$$

Step	Action	Equations
1	**Half-Reactions**	$2\,Cl^- \rightarrow Cl_2 + 2\,e^-$ (oxid) $MnO_2 + 2\,e^- \rightarrow Mn^{2+}$ (red)
2	**Balance atoms: (except H and O)**	**No action necessary**
3	**Balance O: (with H_2O)**	$2\,Cl^- \rightarrow Cl_2 + 2\,e^-$ $MnO_2 + 2\,e^- \rightarrow Mn^{2+} + 2\,H_2O$
4	**Balance H: (with H^+)**	$2\,Cl^- \rightarrow Cl_2 + 2\,e^-$ $4\,H^+ + MnO_2 + 2\,e^- \rightarrow Mn^{2+} + 2\,H_2O$
5	**Balance electrons: (multiply by factor)**	**No action necessary**
6	**Combine and simplify** $4\,H^+ + MnO_2 + 2\,Cl^- \rightarrow Cl_2 + Mn^{2+} + 2\,H_2O$	

b) Oxidation numbers: 2–, +5, +2, +6

$$FeS + NO_3^- \rightarrow NO + SO_4^{2-} + Fe^{2+}$$

Step	Action	Equations
1	**Half-Reactions**	$FeS \rightarrow SO_4^{2-} + Fe^{2+} + 8\,e^-$ (oxid) $NO_3^- + 3\,e^- \rightarrow NO$ (red)
2	**Balance atoms: (except H and O)**	**No action necessary**
3	**Balance O: (with H_2O)**	$4\,H_2O + FeS \rightarrow SO_4^{2-} + Fe^{2+} + 8\,e^-$ $NO_3^- + 3\,e^- \rightarrow NO + 2\,H_2O$
4	**Balance H: (with H^+)**	$4\,H_2O + FeS \rightarrow SO_4^{2-} + Fe^{2+} + 8\,e^- + 8\,H^+$ $4\,H^+ + NO_3^- + 3\,e^- \rightarrow NO + 2\,H_2O$
5	**Balance electrons: (multiply by factor)**	$12\,H_2O + 3\,FeS \rightarrow 3\,SO_4^{2-} + 3Fe^{2+} + 24\,e^- + 24\,H^+$ $32\,H^+ + 8\,NO_3^- + 24\,e^- \rightarrow 8\,NO + 16\,H_2O$
6	**Combine and simplify** $3\,FeS + 8\,NO_3^- + 8\,H^+ \rightarrow 3\,SO_4^{2-} + 3\,Fe^{2+} + 8\,NO + 4\,H_2O$	

2. Balance the following redox reaction in acidic solution, and determine the oxidizing and reducing agents.

0	**+5**	**+2**	**−3**
$Mg\ (s)$ +	$NO_3^{-}\ (aq)$ →	$Mg^{2+}\ (aq)$ +	$NH_4^{+}\ (aq)$
reducing agent	**oxidizing agent**		

Step	Action	Equations
1	**Half-Reactions**	$\mathbf{Mg \rightarrow Mg^{2+} + 2\ e^{-}}$ **(oxid)** $\mathbf{NO_3^{-} + 8\ e^{-} \rightarrow NH_4^{+}}$ **(red)**
2	**Balance atoms: (except H and O)**	$\mathbf{Mg \rightarrow Mg^{2+} + 2\ e^{-}}$ $\mathbf{NO_3^{-} + 8\ e^{-} \rightarrow NH_4^{+}}$
3	**Balance O: (with H_2O)**	$\mathbf{Mg \rightarrow Mg^{2+} + 2\ e^{-}}$ $\mathbf{NO_3^{-} + 8\ e^{-} \rightarrow NH_4^{+} + 3\ H_2O}$
4	**Balance H: (with H^{+})**	$\mathbf{Mg \rightarrow Mg^{2+} + 2\ e^{-}}$ $\mathbf{NO_3^{-} + 10\ H^{+} + 8\ e^{-} \rightarrow NH_4^{+} + 3\ H_2O}$
5	**Balance electrons: (multiply by factor)**	$\mathbf{4\ Mg \rightarrow 4\ Mg^{2+} + 8\ e^{-}}$ $\mathbf{NO_3^{-} + 10\ H^{+} + 8\ e^{-} \rightarrow NH_4^{+} + 3\ H_2O}$
6	**Combine and simplify** $\mathbf{4\ Mg + NO_3^{-} + 10\ H^{+} \rightarrow 4\ Mg^{2+} + NH_4^{+} + 3\ H_2O}$	

3. Balance the following redox reaction in basic solution, and determine the oxidizing and reducing agents.

0	**+5**	**+3**	**−3**
$Al\ (s)$ +	$NO_3^{-}\ (aq)$ →	$Al(OH)_4^{-}$ +	$NH_3\ (aq)$
reducing agent	**oxidizing agent**		

Step	Action	Equations
1	**Half-Reactions**	$\mathbf{Al \rightarrow Al(OH)_4^{-} + 3\ e^{-}}$ **(oxid)** $\mathbf{NO_3^{-} + 8\ e^{-} \rightarrow NH_3}$ **(red)**
2	**Balance atoms: (except H and O)**	$\mathbf{Al \rightarrow Al(OH)_4^{-} + 3\ e^{-}}$ $\mathbf{NO_3^{-} + 8\ e^{-} \rightarrow NH_3}$
3	**Balance O: (with OH^{-} and H_2O)**	$\mathbf{Al + 4\ OH^{-} \rightarrow Al(OH)_4^{-} + 3\ e^{-}}$ $\mathbf{NO_3^{-} + 3\ H_2O\ \ 8\ e^{-} \rightarrow NH_3 + 6\ OH^{-}}$
4	**Balance H: (with H_2O and OH^{-})**	$\mathbf{Al + 4\ OH^{-} \rightarrow Al(OH)_4^{-} + 3\ e^{-}}$ $\mathbf{NO_3^{-} + 6\ H_2O + 8\ e^{-} \rightarrow NH_3 + 9\ OH^{-}}$
5	**Balance electrons: (multiply by factor)**	$\mathbf{8\ Al + 32\ OH^{-} \rightarrow 8\ Al(OH)_4^{-} + 24\ e^{-}}$ $\mathbf{3\ NO_3^{-} + 18\ H_2O + 24\ e^{-} \rightarrow 3\ NH_3 + 27\ OH^{-}}$
6	**Combine and simplify** $\mathbf{8\ Al + 5\ OH^{-} + 3\ NO_3^{-} + 18\ H_2O \rightarrow 8\ Al(OH)_4^{-} + 3\ NH_3}$	

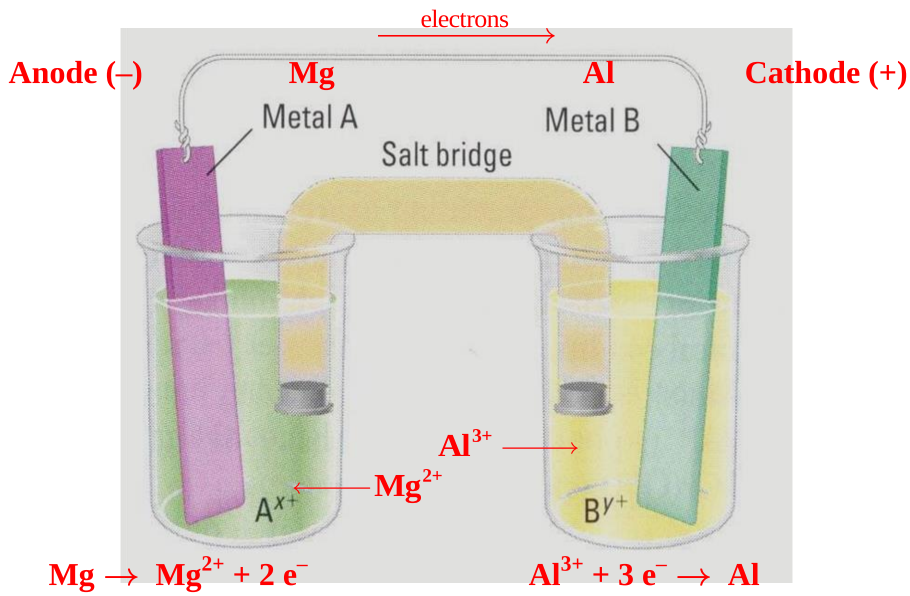
4. Balance the following redox reaction in basic solution, and determine the oxidizing and reducing agents.

$$\overset{+3}{Fe(OH)_3}\ (aq) + O\overset{+1}{Cl^-}\ (aq) \rightarrow \overset{+6}{Fe}O_4^{2-}\ (aq) + \overset{-1}{Cl^-}\ (aq)$$

reducing agent ($Fe(OH)_3$) **oxidizing agent** (OCl^-)

Step	Action	Equations
1	**Half-Reactions**	$Fe(OH)_3 \rightarrow FeO_4^{2-} + 3\,e^-$ **(oxid)** $OCl^- + 2\,e^- \rightarrow Cl^-$ **(red)**
2	**Balance atoms: (except H and O)**	**No action necessary**
3	**Balance O: (with OH^- and H_2O)**	$2\,OH^- + Fe(OH)_3 \rightarrow FeO_4^{2-} + 3\,e^- + H_2O$ $H_2O + OCl^- + 2\,e^- \rightarrow Cl^- + 2\,OH^-$
4	**Balance H: (with H_2O and OH^-)**	$5\,OH^- + Fe(OH)_3 \rightarrow FeO_4^{2-} + 3\,e^- + 4\,H_2O$ $H_2O + OCl^- + 2\,e^- \rightarrow Cl^- + 2\,OH^-$
5	**Balance electrons: (multiply by factor)**	$10\,OH^- + 2\,Fe(OH)_3 \rightarrow 2\,FeO_4^{2-} + 6\,e^- + 8\,H_2O$ $3\,H_2O + 3\,OCl^- + 6\,e^- \rightarrow 3\,Cl^- + 6\,OH^-$
6	**Combine and simplify** $2\,Fe(OH)_3 + 3\,OCl^- + 4\,OH^- \rightarrow 2\,FeO_4^{2-} + 3\,Cl^- + 5\,H_2O$	

5. The diagram below shows a voltaic cell with the anode on the left side and the cathode on the right side. Given that this is a magnesium and aluminum cell,
   - identify metals A and B,
   - identify solutions A and B,
   - write half-reactions for each electrode,
   - direction of electron flow,
   - the polarities of the anode and the cathode,
   - calculate the cell potential, and
   - write a shorthand cell notation.



Anode: $3\ Mg(s) \rightarrow 3\ Mg^{2+}(aq) + 6\ e^-$ $E° = +2.37\ V$

Cathode: $2\ Al^{3+}(aq) + 6\ e^- \rightarrow 2\ Al(s)$ $E° = –1.66\ V$

Overall: $3\ Mg(s) + 2\ Al^{3+}(aq) \rightarrow 3\ Mg^{2+}(aq) + 2\ Al(s)$ $E°_{cell} = 0.71\ V$

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

6. A voltaic cell employs the reaction:

$$Sn\,(s) + 2\,Ag^{+}\,(aq) \rightarrow Sn^{2+}\,(aq) + 2\,Ag\,(s)$$

Calculate the voltage produced by this reaction under standard conditions at 25°C. (Use Table 18.1 in your textbook for standard reduction potentials).

Anode:	$Sn \rightarrow Sn^{2+} + 2\,e^{-}$	**E° = +0.14 volts**
Cathode:	$2\,Ag^{+} + 2\,e^{-} \rightarrow 2\,Ag$	**E° = +0.80 volts**
Overall:	$Sn + 2\,Ag^{+} \rightarrow Sn^{2+} + 2\,Ag$	**$E°_{cell}$ = +0.14 + 0.80 = +0.94 volts**

7. The standard voltage ($\mathcal{E}°$) for the voltaic cell shown below is +0.68 volts.

$$In\,|\,In^{3+}\,||\,Cu^{2+}\,|\,Cu$$

Determine the standard reduction potential for: $In^{3+} + 3\,e^{-} \rightarrow In$

Anode:	$In^{3+} + 3\,e^{-} \rightarrow In$	**E° = ???**
Cathode:	$Cu^{2+} + 2\,e^{-} \rightarrow Cu$	**E° = +0.34 volts**

$E°_{cell} = E°_{cat} - E°_{an}$ = +0.68 volts

$E°_{an} = E°_{cat} - E°_{cell}$ = 0.34 – 0.68 = – 0.34 volts

8. A voltaic cell used the reaction shown below:

$$Sn^{2+}\,(aq) + 2Hg^{2+}\,(aq) \rightarrow Sn^{4+}\,(aq) + Hg_2^{2+}\,(aq)$$

a) Calculate the voltage for this reaction under standard conditions. (Use Table 18.1 in your textbook for standard reduction potentials)

Anode:	$Sn^{2+} \rightarrow Sn^{4+} + 2\,e^{-}$	**E° = –0.15 volts**
Cathode:	$2\,Hg^{2+} + 2\,e^{-} \rightarrow Hg_2^{2+}$	**E° = +0.90 volts**
Overall:	$Sn^{2+} + 2\,Hg^{2+} \rightarrow Sn^{4+} + Hg_2^{2+}$	**E° = +0.90 –0.15= +0.75 volts**

b) Calculate ΔG° for this reaction.

$\Delta G° = w_{max} = -nFE° = -2\text{ mol } e^{-}\,(96500\text{ C/mole } e^{-})(0.75\text{ volts}) = -1.4\times10^{5}$ J

9. A voltaic cell uses the reaction shown below, with a measured standard cell potential of 1.19 V:

$$Tl^{3+}(aq) + 2\,Cr^{2+}(aq) \longrightarrow Tl^{+}(aq) + 2\,Cr^{3+}(aq)$$

a) Write the two half-cell reactions.

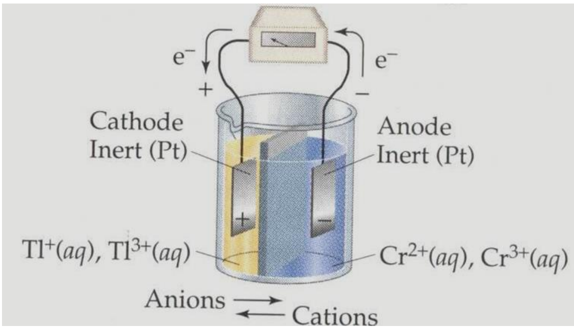
Oxidation	$\mathbf{Cr^{2+}(aq) \longrightarrow Cr^{3+}(aq) + e^{-}}$	**E° = +0.50 V**
Reduction	$\mathbf{Tl^{3+}(aq) + 2\,e^{-} \longrightarrow Tl^{+}(aq)}$	**E° = ???**

b) What is the E° for the reduction of Tl^{3+}?

$E°_{cell} = 0.50 + E°(Tl^{3+}) = 1.19$ V

$E°(Tl^{3+}) = 1.19 - 0.50 =$ 0.69 V

c) Sketch the voltaic cell, label the anode and the cathode, and indicate the direction of the electron flow.



10. For each pair of substances below, use Reduction Potential in your textbook to determine the one that is the stronger oxidizing agent:

The more positive the reduction potential, the easier the substance is reduced and the stronger the oxidizing agent it will be.

a) **$Br_2(l)$** or $I_2(s)$ **(1.09 vs. 0.54)**

b) **$Ag^{+}(aq)$** or $Cu^{+}(aq)$ **(0.80 vs. 0.52)**

c) $Cl_2(g)$ or **$Au^{3+}(aq)$** **(1.36 vs. 1.46)**

d) **Mg(s)** or K(s) **(–2.37 vs. –2.92)**

e) **$H_2O_2(aq)$** or $Cr_2O_7^{2-}(aq)$ **(1.78 vs. 1.33)**

11. The standard cell potential for the reaction shown below is –0.43 V:

$$Eu^{3+}(aq) + e^{-} \longrightarrow Eu^{2+}(aq)$$

Use table of reduction potentials in your textbook to suggest two substances capable of reducing Eu^{3+} to Eu^{2+}.

Any species with a standard reduction potential of less than –0.43 can reduce Eu^{3+} to Eu^{2+}, since it would be a stronger reducing agent that Eu^{3+}. Two examples are Zn and H_2(g).

12. The standard cell potential for the reaction shown below at 298 K is 2.20 V:

$$2\ Al\ (s) + 3\ I_2\ (s) \longrightarrow 2\ Al^{3+}(aq) + 6\ I^{-}\ (aq)$$

Calculate the emf generated by this cell when

$[Al^{3+}] = 4.0 \times 10^{-3} M$
$[I^{-}] = 0.010\ M$

Writing the two half-reactions:

$$2\ Al \rightarrow 2\ Al^{3+} + 6\ e^{-}$$

$$3\ I_2 + 6\ e^{-} \rightarrow 6\ I^{-}$$

Using the Nernst equation:

$$E = E^{\circ} - \frac{0.0592}{n} \log Q \qquad n = 6$$

$$Q = [Al^{3+}]^2\ [I^{-}]^6$$

$$E = 2.20 - \frac{0.0592}{6} \log (4.0 \times 10^{-3})^2 (0.010)^6$$

$$E = 2.20 - (-0.1\underline{6}6) = 2.37$$

13. Use the standard reduction potentials listed in your textbook to determine the equilibrium constant for each of the following reactions:

a) $Zn(s) + Sn^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Sn(s)$

Anode:	$Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$	**E° = +0.76 V**
Cathode:	$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$	**E° = –0.14 V**
Overall:	$Zn(s) + Sn^{2+}(aq) \rightarrow Zn^{2+}(aq) + Sn(s)$	**E° = 0.62 V**

At equilibrium, Nernst equation can be simplified to:

$$0 = E^\circ - \frac{0.0592}{n}\log K$$

$$\log K = \frac{nE^\circ}{0.0592} = \frac{2(0.62)}{0.0592} = 20.95$$

$$K = 10^{20.95} = 8.9 \times 10^{20}$$

b) $Cd(s) + 2\,H^+(aq) \longrightarrow Cd^{2+}(aq) + H_2(g)$

Anode:	$Cd(s) \rightarrow Cd^{2+}(aq) + 2\,e^-$	**E° = +0.40 V**
Cathode:	$2\,H^+(aq) + 2e^- \rightarrow H_2(g)$	**E° = 0.00 V**
Overall:	$Cd(s) + 2\,H^+(aq) \rightarrow Cd^{2+}(aq) + H_2(g)$	**E° = 0.40 V**

$$\log K = \frac{nE^\circ}{0.0592} = \frac{2(0.40)}{0.0592} = 13.5$$

$$K = 10^{13.5} = 3.2 \times 10^{13}$$

14. A voltaic cell utilizes the reaction shown below at 298 K:

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

a) Calculate the emf for this cell under standard conditions.

Anode: $2\,Al(s) \rightarrow 2Al^{3+}(aq) + 6\,e^-$ $E° = +1.66\ V$

Cathode: $3\,Mn^{2+}(aq) + 6e^- \rightarrow 3\,Mn(s)$ $E° = -1.18\ V$

Overall: $2\,Al(s) + 3\,Mn^{2+}(aq) \rightarrow 2\,Al^{3+}(aq) + 3\,Mn(s)$ $E° = 0.48\ V$

b) Calculate the emf for this cell when $[Al^{3+}] = 1.5$ M and $[Mn^{2+}] = 0.10$ M.

$$E = E^\circ - \frac{0.0592}{n}\log Q \qquad n = 6$$

$$E = 0.48 - \frac{0.0592}{6}\log\frac{(1.5)^2}{(0.10)^3}$$

$$E = 0.48 - (0.033) = 0.45\ V$$

15. Use the standard reduction potentials listed in your textbook to determine the equilibrium constant the following reactions:

$$O_2(g) + 4\,H^+(aq) + 4\,Fe^{2+}(aq) \longrightarrow 4\,Fe^{3+}(aq) + 2\,H_2O(l)$$

Anode: $4\,Fe^{2+}(aq) \rightarrow 4\,Fe^{3+}(aq) + 4\,e^-$ $E° = -0.77\ V$

Cathode: $4\,H^+(aq) + O_2(g) + 4e^- \rightarrow 2\,H_2O(l)$ $E° = +1.23\ V$

Overall: $E° = 0.46\ V$

$$\log K = \frac{nE^\circ}{0.0592} = \frac{4(0.46)}{0.0592} = 31.08$$

$$K = 10^{31.08} = 1.2 \times 10^{31}$$