Define each

1. Oxidation - **loss of electrons**
2. Reduction - **gain of electrons**
3. Oxidizing agent - **causes oxidation by undergoing reduction**
4. Reducing agent - **causes reduction by undergoing oxidation**

Write half reactions for each of the following atoms or ions. Label each as oxidation or reduction.

5.	$Al \text{ ----------->} Al^{3+} + 3e^-$	**oxidation**
6.	$S + 2e^- \text{ --------->} S^{2-}$	**reduction**
7.	$2O^{2-} \text{ ---------->} O_2 + 4e^-$	**oxidation**
8.	$Ba^{2+} + 2e^- \text{ ----------->} Ba$	**reduction**
9.	$2N^{3-} \text{ ---------->} N_2 + 6e^-$	**oxidation**
10.	$Br_2 + 2e^- \text{ --------->} 2Br^-$	**reduction**
11.	$P + 3e^- \text{ ---------->} P^{3-}$	**reduction**
12.	$Ca \text{ ----------->} Ca^{2+} + 2e^-$	**oxidation**
13	$Ga^{3+} + 3e^- \text{ ----------->} Ga$	**reduction**
14.	$S + 2e^- \text{ --------->} S^{2-}$	**reduction**
15.	$H_2 \text{ --------->} 2H^+ + 2e^-$	**oxidation**
16.	$2H^+ + 2e^- \text{ --------->} H_2$	**reduction**
17.	$2F^- \text{ ---------->} F_2 + 2e^-$	**oxidation**
18.	$P^{3-} \text{ ---------->} P + 3e^-$	**oxidation**

Balance each spontaneous redox equation. Identify the entities reduced and oxidized. State the reducing agent and the oxidizing agent.

19. Al & Zn^{2+}

2Al +	**$3Zn^{2+}$**	$\rightarrow$	**$2Al^{3+}$** +	**3Zn**
oxidized	**reduced**			
reducing agent	**oxidizing agent**			

20. F_2 & O^{2-}

$2F_2$ + $2O^{2-}$ → $4F^-$ + O_2
reduced **oxidized**
oxidizing agent **reducing agent**

21. O_2 & Ca

2Ca + O_2 → $2Ca^{2+}$ + $2O^{2-}$
oxidized **reduced**
reducing agent **oxidizing agent**

22. Al^{3+} & Li

Al^{3+} + 3Li → Al + $3Li^+$
reduced **oxidized**
oxidizing agent **reducing agent**

Label the species that is **reduced**, that is **oxidized**, the **reducing agent** and the **oxidizing agent**.

23. Fe^{2+} + Co → Co^{2+} + Fe

Co → Co^{2+} + $2e^-$ oxidation **Fe^{2+} + 2e- → Fe reduction**

24. 3 Ag^+ + Ni → Ni^{3+} + 3 Ag

Ni → Ni^{2+} + $2e^-$ oxidation **Ag^+ + 1e- → Agreduction**

25. Cu^{2+} + Pb → Pb^{2+} + Cu

Pb → Pb^{2+} + $2e^-$ oxidation Cu^{2+} + 2e- → Cu reduction

26. O_2 + 2 Sn → O^{2-} + 2 Sn^{2+}

Sn → Sn^{2+} + $2e^-$ oxidation O_2 + 4e- → $2O^{2-}$ reduction

27. Co^{2+} + 2 F^- → Co + F_2

$2F^-$ → F_2 + $2e^-$ oxidation Co^{2+} + 2e- → Coreduction

28. List the species (formulas from above) that lose electrons:

Co Ni Pb Sn F^-

29. List the species (formulas from above) that gain electrons:

Fe^{2+} Ag^+ Cu^{2+} O_2 Co^{2+}

For each of the following reactions, identify:

-The Oxidizing Agent.
-The Reducing Agent.
-The Substance Oxidized.
-The Substance Reduced.

30. I^- + Cl_2 ----------> Cl^- + I_2

Substance oxidized **I^-** Reducing agent **I^-**
Oxidizing agent **Cl_2** Substance reduced **Cl_2**

31. Co + Fe^{3+} -----------> Co^{2+} + Fe^{2+}

Substance oxidized **Co** Reducing agent **Co**
Oxidizing agent **Fe^{3+}** Substance reduced **Fe^{3+}**

32. Cr^{6+} + Fe^{2+} -----------> Cr^{3+} + Fe^{3+}

Substance oxidized **Fe^{2+}** Reducing agent **Fe^{2+}**
Oxidizing agent **Cr^{6+}** Substance reduced **Cr^{6+}**

Redox Half Reactions and Reactions WS #2

1. State the Oxidation Number of each of the elements that is underlined.

a) $\underline{N}H_3$	**-3**	b) $H_2\underline{S}O_4$	**6**
c) $Zn\underline{S}O_3$	**4**	d) $\underline{Al}(OH)_3$	**3**
e) $\underline{Na}$	**0**	f) $\underline{Cl}_2$	**0**
g) $Ag\underline{N}O_3$	**5**	h) $\underline{Cl}O_4^-$	**7**
i) $\underline{S}O_2$	**4**	j) $K_2\underline{Cr}_2O_4$	**3**
k) $Ca(\underline{Cl}O_3)_2$	**5**	l) $K_2\underline{Cr}_2O_7$	**6**
m) $H\underline{P}O_3^{2-}$	**3**	n) $H\underline{Cl}O$	**1**
o) $\underline{Mn}O_2$	**4**	p) $K\underline{Cl}O_3$	**5**
q) $\underline{Pb}O_2$	**4**	r) $\underline{Pb}SO_4$	**2**
s) $K_2\underline{S}O_4$	**6**	t) $\underline{N}H_4^+$	**-3**
u) $Na_2\underline{O}_2$	**-1**	v) $\underline{Fe}O$	**2**
w) $\underline{Fe}_2O_3$	**3**	x) $Si\underline{O}_4^{4-}$	**-2**
y) $Na\underline{I}O_3$	**5**	z) $\underline{Cl}O_3^-$	**5**

aa) $\underline{N}O_3^-$ **5** bb) $\underline{Cr}(OH)_4$ **4**

cc) $\underline{Ca}H_2$ **-1** dd) $\underline{Pt}(H_20)_5(0H)^{2+}$ **+3**

ee) $\underline{Fe}(H_2O)_6^{3+}$ **+3** ff) $\underline{C}H_3COOH$ **0**

2. What is the oxidation number of carbon in each of the following substances?
   a) CO **2** b) C **0**
   c) CO_2 **4** d) CO_3^{2-} **4**
   e) C_2H_6 **-3** f) CH_3OH **-2**

3. For each of the following reactions, identify: the oxidizing agent, the reducing agent, the substance oxidized and the substance reduced.

a) $Cu^{2+}{}_{(aq)}$ + $Zn_{(s)}$ --------> $Cu_{(s)}$ + $Zn^{2+}{}_{(aq)}$

Substance oxidized **Zn** Substance reduced **Cu^{2+}**

Oxidizing agent **Cu^{2+}** Reducing agent **Zn**

b) $Cl_{2\,(g)}$ + 2 $Na_{(s)}$ --------> 2 $Na+_{(aq)}$ + 2 $Cl^-{}_{(aq)}$

Substance oxidized **Na** Substance reduced **Cl_2**

Oxidizing agent **Cl_2** Reducing agent **Na**

WS # 3 Spontaneous and Non-spontaneous Redox Reactions

Describe each reaction as spontaneous or non-spontaneous.

1. Au^{+3} + Fe^{+3} -----> Fe^{+2} + Au **nonspontaneous (two oxidizing agents)**

2. Pb + Fe^{+3} ------> Fe^{+2} + Pb^{+2} **spontaneous**

3. Cl_2 + F^- ------> F_2 + $2Cl^-$ **nonspontaneous**

4. $S_2O_8^{-2}$ + Pb ------> $2SO_4^{-2}$ + Pb^{+2} **spontaneous**

5. Cu^{+2} + $2Br^-$ ------> Cu + Br_2 **nonspontaneous**

6. Sn^{+2} + Br_2 ------> Sn^{+4} + $2Br^-$ **spontaneous**

7. Pb^{+2} + Fe^{+2} ------> Fe^{+3} + Pb **nonspontaneous**

8. Can you keep 1 M HCl in an iron container. If the answer is no, write a balanced equation for the reaction that would occur. **No**

Fe + $2H^{+}$ --------> Fe^{2+} + H_2

9. Can you keep 1 M HCl in an Ag container. If the answer is no, write a balanced equation for the reaction that would occur.

Yes. There is no reaction.

10. Can you keep 1 M HNO_3 in an Ag container. If the answer is no, write a balanced equation for the reaction that would occur. (remember HNO_3 consists of two ions H^{+} and NO_3^{-})

No 3Ag + NO_3^{-} + $4H^{+}$ --------> $3Ag^{+}$ + NO + $2H_2O$

11. Can you keep 1 M HNO_3 in an Au container. If the answer is no, write a balanced equation for the reaction that would occur. (Remember, HNO_3 consists of two ions H^{+} and NO_3^{-})

Yes. There is no reaction.

12. Circle each formula that is able to lose an elecron

O_2 **Cl^{-}** **Fe** Na^{+}

13. Determine the oxidation number for the element underlined.

$Pb\underline{S}O_4$	**6**	$\underline{Cl}O_3^{-}$	**5**
$H\underline{P}0_3^{2-}$	**3**	$Na_2\underline{O}_2$	**-1**
$Ca\underline{H}_2$	**-1**	$Al_2(\underline{S}O_4)_3$	**6**
$Na\underline{I}O_3$	**5**	$\underline{C}_4H_{12}$	**-3**

14. Al^{3+} + Zn ---------> Al + Zn^{2+}

Substance oxidized **Zn** Oxidizing agent **Al^{3+}**

15. $Cr_2O_7^{2-}$ + ClO_2^- ------------> Cr^{3+} + ClO_4^-

Substance reduced **$Cr_2O_7^{2-}$** Oxidizing agent **$Cr_2O_7^{2-}$**

16. State the Oxidation Number of each of the elements that is underlined.

a) $\underline{N}H_3$	**-3**	b) $H_2\underline{S}O_4$	**6**
c) $Zn\underline{C}O_3$	**4**	d) $\underline{Al}(OH)_3$	**3**
e) $\underline{Na}$	**0**	f) $\underline{Cl}_2$	**0**

17. Balance the redox equation using the half reaction method.

Al + $3Ag^+$ ----------> Al^{3+} + 3Ag

18. Circle each formula that is able to lose an electron

O_2 **Cl^-** **Fe** Na^+

Determine the oxidation number for the element underlined.

19. $\underline{Pb}SO_4$ **2**

20. $\underline{Cl}O_3^-$ **5**

21. $H\underline{P}O_3^{2-}$ **3**

22. $Na_2\underline{0}_2$ **-1**

23. $Ca\underline{H}_2$ **-1**

24. $Na\underline{I}O_3$ **5**

25. $\underline{C}_4H_{12}$ **-3**

26. $Al_2(\underline{S}O_4)_3$ **6**

27. Al^{3+} + Zn ----------> Al + Zn^{2+}

Substance oxidized **Zn** Oxidizing agent **Al^{3+}**

28. $Cr_2O_7^{2-}$ + ClO_2^- ----------------> Cr^{3+} + ClO_4^-

Substance reduced $Cr_2O_7^{2-}$ Oxidizing agent
$Cr_2O_7^{2-}$

29. $O_3 + H_2O + SO_2 \text{----->} SO_4^{2-} + O_2 + 2H^+$

Substance oxidized SO_2 Reducing agent SO_2

30. $3As_2O_3 + 4NO_3^- + 7H_2O + 4H^+ \text{-------->} 6H_3AsO_4 + 4NO$

Substance reduced NO_3^- Reducing agent
As_2O_3

WS # 4 Balancing Redox Reactions

Balance each of the following half-cell reactions. (In each case assume that the reaction takes place in an **ACIDIC** solution.) Also, state whether the reaction is oxidation or reduction.

1. $5H_2O + S_2O_3^{2-} \text{-------------->} 2SO_4^{2-} + 10H^+ + 8e^-$

   oxidation

2. $8H^+ + 5e^- + MnO_4^- \text{-------------->} Mn^{2+} + 4H_2O$

   reduction

3. $4H_2O + As \text{-------------->} AsO_4^{3-} + 8H^+ + 5e^-$

   oxidation

4. $7H_2O + 2Cr^{3+} \text{----------->} Cr_2O_7^{2-} + 14H^+ + 6e^-$

   oxidation

5. $2H_2O + Pb^{2+} \text{-------------->} PbO_2 + 4H^+ + 2e^-$

   oxidation

6. $8H^+ + SO_4^{2-} + 6e^- \text{--------------->} S + 4H_2O$

reduction

7. $4H^+ + NO_3^- + 3e^- \text{-------------->} NO + 2H_2O$

reduction

8. $10H^+ + 8e^- + NO_3^- \text{-------------->} NH_4^+ + 3H_2O$

reduction

9. $12H^+ + 10e^- + 2BrO_3^- \text{--------------->} Br_2 + 6H_2O$

reduction

Balancing Half Cell Reactions

Balance in basic solution.

10. $3e^- + 2H_2O + NO_3^- \text{--------------->} NO + 4OH^-$

11. $4H_2O + 5e^- + MnO_4^- \text{--------------->} Mn^{2+} + 8OH^-$

12. $8OH^- + As \text{--------------->} AsO_4^{3-} + 4H_2O + 5e^-$

13. $14OH^- + 2Cr^{3+} \text{--------------->} Cr_2O_7^{2-} + 7H_2O + 6e^-$

14. $4OH^- + Pb^{2+} \text{--------------->} PbO_2 + 2H_2O + 2e^-$

15. $4H_2O + 6e^- + SO_4^{2-} \text{--------------->} S + 8OH^-$

16. $10\ OH^- + S_2O_3^{2-} \text{--------------->} 2SO_4^{2-} + 5H_2O + 8e^-$

17. $7H_2O + 8e^- + NO_3^- \text{--------------->} NH_4^+ + 10\ OH^-$

18. **$6H_2O$ + $10e^-$ + $2BrO_3^-$ --------------> Br_2 + $12\ OH^-$**

19. Determine if each of the following changes is oxidation, reduction or neither.

SO_3^{2-}	-------->	SO_4^{2-}	**oxidation**
CaO	-------->	Ca	**reduction**
CrO_4^{2-}	-------->	$Cr_2O_7^{2-}$	**neither**
CrO_4^{2-}	-------->	Cr^{3+}	**reduction**
$2I^-$	-------->	I_2	**oxidation**
IO_3^-	-------->	I_2	**reduction**
MnO_4^-	-------->	Mn^{2+}	**reduction**
ClO_2^-	-------->	ClO^-	**reduction**

20. $Cr_2O_7^{2-}$ + Fe^{2+} --------> Cr^{3+} + Fe^{3+}

Substance oxidized **Fe^{2+}** Substance reduced **$Cr_2O_7^{2-}$**

Oxidizing agent **$Cr_2O_7^{2-}$** Reducing agent **Fe^{2+}**

WS #5 Balancing Redox Reactions in Acid and Basic Solution

Balance each redox equation. Assume all are spontaneous. Use the half reaction method.

1. $2O^{2-} + 2F_2 \text{ -----------> } O_2 + 4F^-$

2. $4Al + 3O_2 \text{ -----------> } 6O^{2-} + 4Al^{3+}$

3. $2K + Zn^{+2} \text{ -----------> } Zn + 2K^+$

Balance each half reaction in basic solution.

4. $Cr_2O_7^{2-} + 7H_2O + 6e^- \text{ --------------> } 14OH^- + 2Cr^{3+}$

5. $NO + 4OH^- \text{ -----------------> } 2H_2O + NO_3^- + 3e^-$

6. $2H_2O + 2e- + SO_4^{2-} \text{ -------------> } SO_2 + 4OH^-$

7. $2MnO_2 + H_2O + 2e- \text{ -------------> } Mn_2O_3 + 2OH^-$

Balance each redox reaction in <u>acid</u> solution using the half reaction method.

8. $8H^+ + 3H_2O_2 + Cr_2O_7^{2-} \text{ -------> } 3O_2 + 2Cr^{3+} + 7H_2O$

9. $TeO_3^{2-} + 2N_2O_4 + H_2O \text{ -------> } Te + 4NO_3^- + 2H^+$

10. $4H^+ + 4ReO_4^- + 7IO^- \text{ -------> } 7IO_3^- + 4Re + 2H_2O$

11. $8H^+ + 5PbO_2 + I_2 \text{ -------> } 5Pb^{2+} + 2IO_3^- + 4H_2O$

12. $12H_2O + 8As \text{ -------> } 3H_2AsO_4^- + 5AsH_3 + 3H^+$

Balance each redox reaction in basic solution using the half reaction method.

13. $3O_2 + 8OH^- + 2Cr^{3+}$ -------> $H_2O + 3H_2O_2 + Cr_2O_7^{2-}$

14. $H_2O + Te + 4NO_3^-$ ------->$TeO_3^{2-} + 2OH^- + 2N_2O_4$

15. $7IO_3^- + 4OH^- + 4Re$ -------> $4ReO_4^- + 7IO^- + 2H_2O$

16. $8OH^- + 5Pb^{2+} + 2IO_3^-$ -------> $5PbO_2 + I_2 + 4H_2O$

17. $7H_2O + Cr_2O_7^{2-} + 3Hg$ -------> $3Hg^{2+} + 14OH^- + 2Cr^{3+}$

State of the change represents oxidation, reduction or neither (use oxidation #s).

18.	MnO_2	-------->	Mn_2O_3		**reduction**
19.	NH_3	-------->	NO_2		**oxidation**
20.	$HClO_4$	------->	$HCl + H_2O$		**reduction**
21.	O_2	-------->	O^{2-}		**reduction**
22.	P_2O_5	-------->	P_4H_{10}		**reduction**

Determine the oxidation number

23.	$H_2\underline{S}O_4$	**6**	22.	$H\underline{S}O_4^-$	**6**
24.	$\underline{P}_4$	**0**	23.	$Na\underline{H}$	**-1**
25.	$\underline{U}O_3$	**6**	24.	$Na_2\underline{O}_2$	**-1**
26.	$\underline{U}_2O_5$	**5**	25.	$\underline{Pb}SO_4$	**2**

WS #6 Review

1. Describe each in your own words

1. Oxidation - **loss of electrons**
2. Reduction - **gain of electrons**
3. Oxidizing agent - **causes oxidation by undergoing reduction**
4. Reducing agent - **causes reduction by undergoing oxidation**

2. Write half reactions for each. Describe as oxidation or reduction. Circle all oxidizing agents.

a)	$Na \longrightarrow Na^{+} + e^{-}$	**oxidation**
b)	$Ca \longrightarrow Ca^{2+} + 2e^{-}$	**oxidation**
c)	$Al^{3+} + 3e^{-} \longrightarrow Al$	**reduction**
d)	$2F^{1-} \longrightarrow F_2 + 2e^{-}$	**oxidation**
e)	$N_2 + 6e^{-} \longrightarrow 2N^{3-}$	**reduction**
f)	$2O^{2-} \longrightarrow O_2 + 4e^{-}$	**oxidation**

3. Write the reaction between the following: Use the half reaction method.

a) $Ca + Al(NO_3)_3$

$3Ca + 2Al^{3+} \longrightarrow 2Al + 3Ca^{2+}$

b) $Sn + AgNO_3$

$Sn + 2Ag^{+} \longrightarrow 2Ag + Sn^{2+}$

c) $Sn + Au(NO_3)_3$

$3Sn + 2Au^{3+} \longrightarrow 2Au + 3Sn^{2+}$

4. Circle each reducing agent: **Cu** Cu^+ **Al** Al^{3+}

5. Circle each oxidizing agent: F^- **F** O^{2-} **O_2**

6. Ni^{+2} reacts with Mn, however, Al^{+3} does not react with Mn. Rank the oxidizing agents in order of decreasing strength. Rank the reducing agents in order of decreasing strength.

strongest oxidizing agent $Ni^{2+} + 2e^{-} \longrightarrow Ni$

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

Al^{3+} + $3e^-$ -----------> Al strongest reducing agent

7. Ag^+ reacts with Pb, however, Ca^{+2} does not react with Pb. Rank the reducing agents in order of decreasing strength. Rank the oxidizing agents in order of decreasing strength.

strongest oxidizing agent Ag^+ + $1e^-$ -----------> Ag
Pb^{2+} + $2e^-$ -----------> Pb
Ca^{2+} + $2e^-$ -----------> Ca strongest reducing agent

8. Cl_2 reacts with Ag, however, Ag does not react with Mg^{+2}. Rank the oxidizing agents in order of decreasing strength. Rank the reducing agents in order of decreasing strength.

strongest oxidizing agent Cl_2 + $2e^-$ --------> $2Cl^-$
Ag^+ + $1e^-$ -----------> Ag
Mg^{2+} + $2e^-$ -----------> Mg strongest reducing agent

9. Ni^{+2} reacts with Mn, however, Al^{+3} does not react with Mn. Rank the reducing agents in order of decreasing strength. Rank the oxidizing agents in order of decreasing strength.

strongest oxidizing agent Ni^{2+} + $2e^-$ -----------> Ni
Mn^{2+} + $2e^-$ -----------> Mn
Al^{3+} + $3e^-$ -----------> Al strongest reducing agent

10. Cl_2 reacts with Br^-, however, I_2 does not react with Br^-. Rank the oxidizing agents in order of decreasing strength. Rank the reducing agents in order of decreasing strength.

strongest oxidizing agent Cl_2 + $2e^-$ --------> $2Cl^-$
Br_2 + $2e^-$ --------> $2Br^-$
I_2 + $2e^-$ --------> $2I^-$ strongest reducing agent

Classify as oxidation, reduction or neither.

11. SO_4^{2-} --------> S^{2-} **reduction**
12. MnO_2 --------> MnO_4^- **oxidation**
13. $Cr_2O_7^{2-}$ --------> CrO_4^{2-} **neither**
14. IO_3^- --------> I_2 **reduction**

15. Given the following lab data

$SnCl_2$	&	Ni	Spontaneous
$Ni(NO_3)_2$	&	Fe	Spontaneous
$Cr(NO_3)_3$	&	Fe	Non spontaneous.

i) Write three balanced equations.

Ni + Sn^{2+} -------------> Ni^{2+} + Sn

Fe + Ni^{2+} -------------> Fe^{2+} + Ni

Fe + Cr^{3+} <------------- Fe^{2+} + Cr

ii) Rank the oxidizing agents in decreasing order of strength.

strongest oxidizing agent **Sn^{2+} + $2e^-$ -----------> Sn**

Ni^{2+} + $2e^-$ -----------> Ni

Fe^{2+} + $2e^-$ -----------> Fe

Cr^{3+} + $3e^-$ -----------> Cr **strongest reducing agent**

iii) Rank the reducing agents in decreasing order of strength. **See above.**

iv) Will $SnCl_2$ react with Cr? Explain? **Yes, because Sn^{2+} is a stronger oxidizing agent than Cr^{3+} .**

v) Will Fe^{2+} react with Sn? **No, because Fe^{2+} is a weaker oxidizing agent than Sn^{2+}**

16. **$2H^+$ + $2MnO_4^-$ + $5H_2S$ --------> 5S + $6H_2O$ + 2MnO**

oxidizing agent **reducing agent**

17. **$2H^+$ + $10SO_4^{2-}$ + $4Br_2$ ----------> $5S_2O_3^{2-}$ + $8BrO_3^-$ + H_2O**

oxidizing agent **reducing agent**

18. Balance in basic solution

$2MnO_4^-$ + $5H_2S$ --------> 5S + 2MnO + $4H_2O$ + $2OH^-$

19. Describe as spontaneous or non-spontaneous. Use your reduction potential chart.

a) $ZnCl_2$ & Cu **nonspontaneous**

b) $CuCl_2$ & NaCl **nonspontaneous**
c) Br_2 & Fe^{2+} **spontaneous**
d) H_2S & Al^{3+} **nonspontaneous**

20. Can you keep HCl in a Zn container? **No, Spontaneous reaction.**

What about an Au container? **Yes, nonspontaneous reaction.**

Balance in basic solution

21. **H_2O + $10SO_4^{2-}$ + $4Br_2$ ------> $5S_2O_3^{2-}$ + $2OH^-$ + $8BrO_3^-$**

Classify as an oxidizing agent, reducing agent or both based on its position on the table. State the E^o or voltage of its position. Some of these are both, so state two voltages and indicate that it can be an oxidizing and reducing agent.

e.g.	MnO_4^- (in acid)	**oxidizing agent**	**1.51 v**
22.	Br_2	**oxidizing agent**	**1.09 v**
23.	Fe^{2+}	**oxidizing agent / reducing agent**	**-0.45 v / 0.77 v**
24.	MnO_4^- (water)	**oxidizing agent**	**0.60 v**
25.	Ni	**reducing agent**	**-0.26 v**
26.	Cr^{3+}	**oxidizing agent**	**-0.74 v**
27.	H_2O	**oxidizing agent / reducing agent**	**-0.40 v / +0.80 v**

Indicate as spontaneous or non-spontaneous.

28.	MnO_4^-	&	Fe^{2+}	**non-spontaneous**
29.	Cu^{2+}	&	Br^-	**non-spontaneous**
30.	HNO_3	&	Ag	**spontaneous**
31.	MnO_4^- (acid)	&	H_2O	**spontaneous**
32.	$Ni_{(s)}$	&	Al^{3+}	**non-spontaneous**
33.	HCl	&	Mg	**spontaneous**

Write each oxidation and reduction half reaction for each question above. Determine the E^o for each. Calculate the E^o for the overall reaction.

34. $MnO_4^- + 2H_2O + 3e^- \text{-------->} MnO_2 + 4OH^-$
+0.60 v
$3(Fe^{2+} \text{---------->} Fe^{3+} + 1e^-)$
-0.77 v
$MnO_4^- + 2H_2O + 3Fe^{2+} \text{---------->} 3Fe^{3+} + MnO_2 + 4OH^-$
-0.17 v

35.

36. $NO_3^- + 4H^+ + 3e^- \text{---------->} NO + 2H_2O$
+0.96 v
$3(Ag \text{---------->} Ag^+ + 1e^-)$
-0.80 v
$NO_3^- + 4H^+ + 3Ag \text{---------->} NO + 2H_2O + 3Ag+$
+0.16 v

37.

38.

39. $2H^+ + 2e^- \text{------>} H_2$
0.00 v
$Mg \text{---------->} Mg^{2+} + 2e^-$
2.37 v
$Mg + 2H^+ \text{---------->} Mg^{2+} + H_2$
2.37 v

WS # 7 Electrochemical Cells

1. Oxidation is when electrons are **lost**.

2. Reduction is when electrons are **gained**.

3. The reducing agent undergoes **oxidation**.

4. The oxidizing agent undergoes **reduction**.

5. A negative voltage means the reaction is **nonspontaneous**.

6. In an electrochemical cell electrons exit the electrode, which is **negative**.

7. In an electrochemical cell the reduction reaction is **higher** on the chart, while the oxidation reaction is **lower.** .

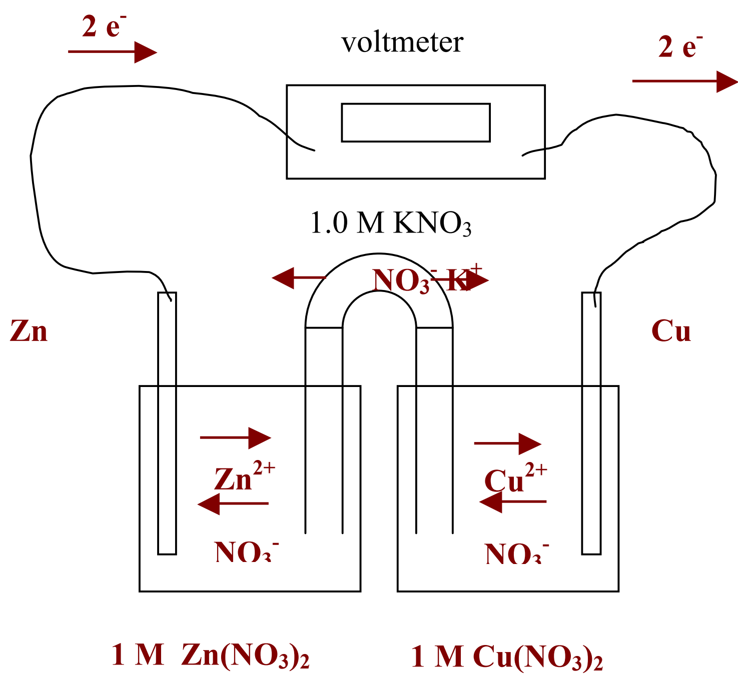
8. The cathode is the site of **reduction** and the anode is the site of **oxidation.**
.

9. Anions migrate to the **anode** and cations migrate to the **cathode**.

10. Anions have a **negative** charge and cations have a **positive** charge.

Draw and completely analyze each electrochemical cell.

11. $Zn / Zn(NO_3)_2 \parallel Cu / Cu(NO_3)_2$

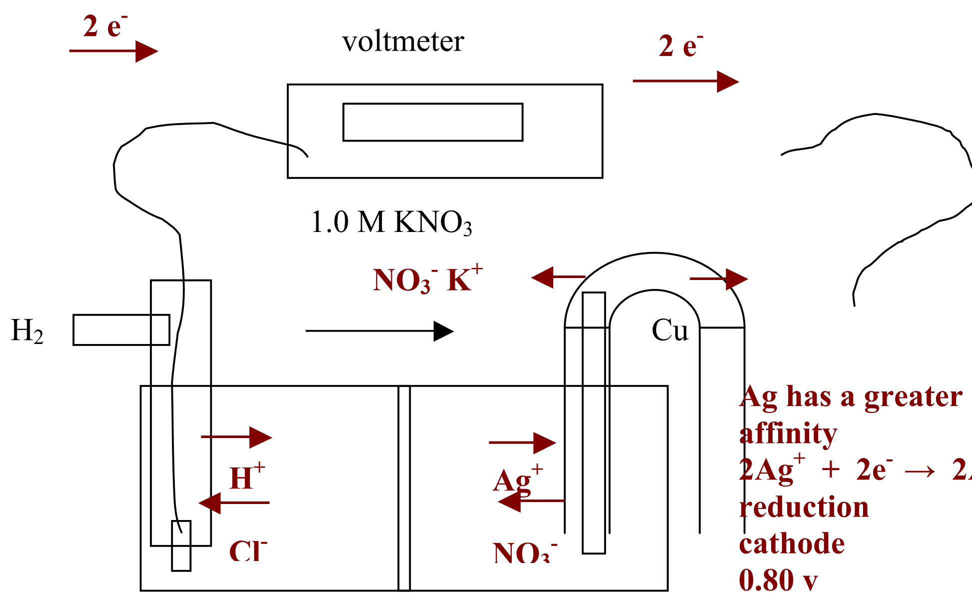


$Zn \rightarrow Zn^{2+} + 2e^-$
oxidation
anode
0.76 v
loses mass

Cu has greater electron affinity
$Cu^{2+} + 2e^- \rightarrow Cu$
reduction
cathode
0.34 v
gains mass

$Cu^{2+} + Zn \rightarrow Zn^{2+} + Cu$ 1.10 v

12. Ag / $AgNO_3$ ‖ H_2 / HCl



$H_2 \rightarrow 2H^+ + 2e^-$
oxidation
anode
0.00 v

Ag has a greater electron affinity
$2Ag^+ + 2e^- \rightarrow 2Ag$
reduction
cathode
0.80 v

1 M HCl **1 M $Ag(NO_3)_2$**

$2Ag^+ + H_2 \rightarrow 2Ag + 2H^+$ 0.80 v

WS # 8

1. In an electrolytic cell, reduction occurs at the **negative** electrode and oxidation occurs at the **positive** electrode.

2. If there are two possible reduction reactions, the **highest** one on the chart occurs.

3. For reduction, the chart is read from **left** to **right**.

4. For oxidation, the chart is read from **right** to **left** and the sign of the voltage is **changed.**

5. If there are two possible oxidation reactions, the **lowest** one on the chart occurs.

6. Corrosion of a metal is **oxidation.**

7. Electrolysis **uses** electrical energy.

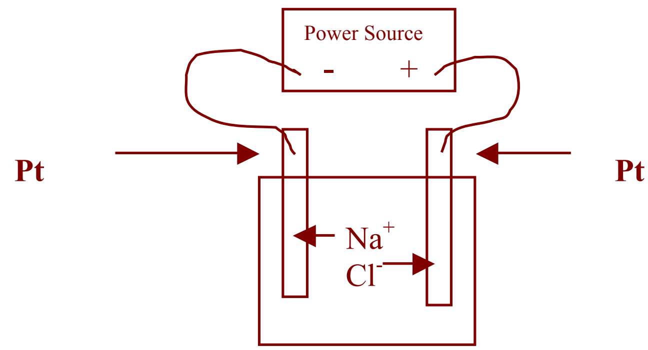
8. Electrochemical cells **produce** electrical energy.

9. Electrolytic cells **use** electrical energy.

10. What is the standard reference cell? **hydrogen** E^o = **O** v

Draw and completely analyze each electrolytic cell.

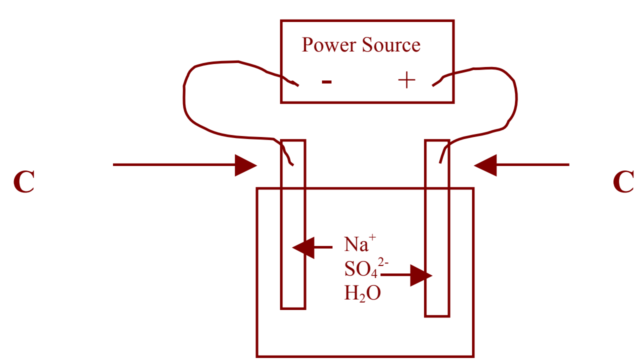
11. Molten NaCl



Cathode: $Na^+ + 1e^- \rightarrow Na_{(s)}$ -2.71 v **Anode: $2Cl^- \rightarrow Cl_2 + 2e^-$ -1.36 v**

Overall: $2Na^+ + 2Cl^- \rightarrow Cl_2 + 2Na_{(s)}$ -4.07 v **MTV = +4.07 v**

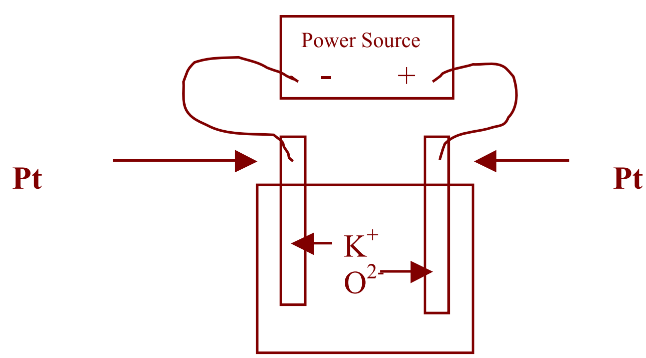
12. Aqueous Na_2SO_4



Cathode: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ -0.41 v **Anode: $H_2O \rightarrow 2H^+ + 1/2O_2 + 2e^-$ -0.82 v**

Overall: $H_2O \rightarrow H_2 + 1/2O_2$ -1.23 v **MTV = +1.23 v**

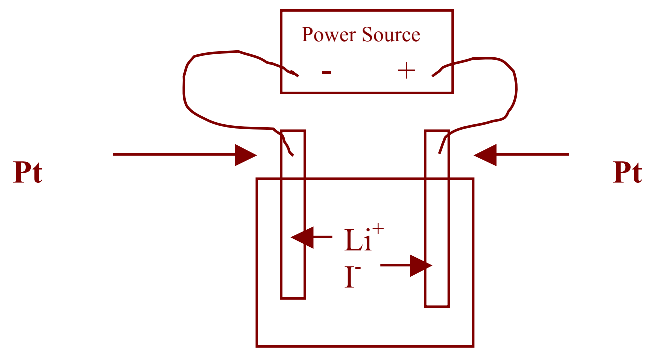
13. Liquid K_2O



Cathode: $K^+ + 1e^- \rightarrow K_{(s)}$ -2.93 v **Anode: $2O^{2-} \rightarrow O_2 + 4e^-$? v**

Overall: $4K^+ + 2O^{2-} \rightarrow O_2 + 4K_{(s)}$ -? v   MTV = +? v

14. 1.0 M LiI



Cathode: Cathode: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ -0.41 v   Anode: $2I^- \rightarrow I_2 + 2e^-$ -0.54 v

Overall: $2H_2O + 2I^- \rightarrow I_2 + H_2 + 2OH^-$ -0.95 v   MTV = +0.95 v

15. 250ml of 0.200M MnO_4^- reacts with excess SO_3^{-2}. How many grams of MnO_2 are produced? This is Chemistry 11 stoichiometry. $2MnO_4^- + 3SO_3^{-2} + H_2O$ -----> $2MnO_2 + 3SO_4^{-2} + 2OH^-$

$0.250L\ MnO_4^- \times \frac{0.200\ mol}{L} \times \frac{2\ mol\ MnO_2}{2\ mol\ MnO_4^-} \times \frac{86.9g}{mol} = 4.34g$

16. Determine the oxidation number for each underlined atom.

$\underline{Mn}O_2$ **4** $\underline{Cr}_2O_7^{-2}$ **6** $\underline{I}O_3^-$ **5** $\underline{C}_2O_4^{-2}$ **3** $Al(\underline{N}O_3)_3$ **5**

17. Describe each term:

Salt bridge- **a u-tube filled with salt solution that allows ions to flow in an electrochemical cell.**
Electrolyte- **a solution that conducts electricity**
Anode- **an electrode that is the site of oxidation**
Cathode- **an electrode that is the site of reduction**
Spontaneous- **a reaction that occurs naturally and has a positive voltage**
Electron affinity- **the ability of a metal to attract electrons**

18. What would happen if you used an aluminum spoon to stir a solution of $FeSO_{4(aq)}$? Write a reaction and calculate E^o.

$2Al + 3Fe^{2+} \text{------->} 2Al^{3+} + 3Fe$ $\quad E^0 = 1.21$ v Spontaneous. There would be a reaction!

19. Draw an electrochemical cell using Cu and Ag electrodes.

Cathode (+)	**Anode (-)**
Ag	**Cu**
$Ag^+ + 1e^-$--------> Ag 0.80v	**Cu -------> Cu^2 + 2e**
-0.34v	

$2Ag^+ + Cu \text{------>} 2Ag + Cu^{2+}$ $\quad E^0 = 0.46$ v spontaneous

20. 250ml of .500M MnO_4^- are required to titrate a 100ml sample of SO_3^{-2}. Calculate the $[SO_3^{-2}]$

$$2MnO_4^- + 3SO_3^{-2} + H_2O \text{----->} 2MnO_2 + 3SO_4^{-2} + 2OH^-$$

$$\frac{.250L\ MnO_4^- \times \frac{0.500\ mol}{L} \times \frac{3\ mol\ SO_3^{-2}}{2MnO_4^-}}{0.100L} = 1.88M$$

21. How is the breathalyzer reaction used to determine blood alcohol content (you might need to look this up in your textbook)?

The breathalyzer reaction uses a spontaneous redox reaction between acidic $Cr_2O_7^{2-}$ and ethanol C_2H_5OH. If alcohol is present in your breath sample, it will react with a solution of $Cr_2O_7^{2-}$ reducing the orange color as it reacts to form Cr^{3+}, which is green. The drunker you are, the greater the reduction in orange color, which is measured with a spectrophotometer.

22. $2H^+ + Mg \text{----->} Mg^{+2} + H_2$

Oxidizing agent **H^+** Reducing agent **Mg**

WS #9 Electrolytic, Electrochemical Cells & Application

Determine the half reactions for each cell and the cell voltage or minimum theoretical voltage and overall equation.

1. Ag / Pb electrochemical cell.

Anode: **Pb** **Cathode:** **Ag**
Anode reaction: **$Pb \text{-------->} Pb^{2+} + 2e^-$** Cathode reaction: **$Ag^+ + 1e^- \text{------->} Ag$**
Overall reaction: **$Pb + 2Ag+ \text{----->} Pb^{2+} + 2Ag$** Voltage: **0.93v**

2. $ZnCl_{2(l)}$ electrolytic cell (electro-winning)

Anode: **C** **Cathode:** **C**
Anode reaction: **$2Cl^- \text{-------->} Cl_2 + 2e^-$** Cathode reaction: **$Zn^{2+} + 2e^- \text{------->} Zn$**
Overall reaction: **$2Cl^- + Zn^{2+} \text{----->} Cl_2 + Zn$** MTV: **+2.12 v**

3. $CuSO_{4(aq)}$ electrolytic cell (electro-winning)

Anode: **C** **Cathode:** **C**
Anode reaction: **$H_2O \text{-------->} 2H^+ + 1/2O_2 + 2e^-$** Cathode reaction: **$Cu^{2+} + 2e^- \text{------->} Cu$**
Overall reaction: **$H_2O + Cu^{2+} \text{----->} 2H^+ + 1/2O_2 + Cu$** MTV: **+0.48 v**

4. The electrolysis of 1M NaI (electro-winning)

Anode: **C** **Cathode:** **C**
Anode reaction: **$2I^- \text{-------->} I_2 + 2e^-$** Cathode reaction: **$2H_2O + 2e^- \text{------->} H_2 + 2OH^-$**

Overall reaction: $2H_2O$ + $2I^-$ -----> H_2 + $2OH^-$ + I_2 MTV: **+0.95 v**

5. The reaction needed to make Al. The electrolyte is **Al_2O_3** and its phase is **molten** (molten or aqueous).

To lower the mp. from 2000 ^oC to 800 ^oC **cryolite** is used.

Anode: **C** **Cathode:** **C**

Anode reaction: $2O^{2-}$ -------> O_2 + $4e^-$ Cathode reaction: Al^{3+} + $3e^-$ -------> Al

Overall reaction: $6O^{2-}$ + $4Al^{3+}$ -----> $3O_2$ + 4Al

6. The reaction needed to electroplate a copper penny with silver.

Anode: **Ag** Cathode: **penny**

Anode reaction: Ag-----> Ag^+ + e^- Cathode reaction: Ag^+ + e^- -----> Ag

7. The reaction needed to **nickel plate** a copper penny.

Anode: **Ni** Cathode: **penny**

Anode reaction: Ni-----> Ni^{+2} + $2e^-$ Cathode reaction: Ni^{2+} + $2e^-$ -----> Ni

Possible Electrolyte **$Ni(NO_3)_2$**

8. The reaction used in the **electrorefining** of lead.

Anode: **Impure Lead** Cathode: **Pure Lead**

Anode reaction: Pb-----> Pb^{+2} + $2e^-$ Cathode reaction: Pb^{2+} + $2e^-$ -----> Pb

WS # 10 Electrolytic, Electrochemical Cells, Corrosion, & Cathodic Protection

Determine the half reactions for each cell and the cell voltage or minimum theoretical voltage.

1. Zn / Mg electrochemical cell

Anode: **Mg** **Cathode:** **Zn**

Anode reaction: **Mg --------> Mg^{2+} + $2e^-$** Cathode reaction: **Zn^{+2} + $2e^-$ -------> Zn**
Overall reaction: **Mg + Zn^{2+}-----> Mg^{2+} + Zn** Voltage: **1.61v**

2. The electrolytic cell used to produce Al.

Electrolyte: **Al_2O_3** Phase (aqueous or molten) **Molten**
Anode: C **Cathode: C**
Anode reaction: **$2O^{2-}$ -------> O_2 + $4e^-$** Cathode reaction: **Al^{3+} + $3e^-$ -------> Al**
Overall reaction: **$6O^{2-}$ + $4Al^{3+}$ -----> $3O_2$ + $4Al$**

3. The electrolysis KI(aq)

Anode: C **Cathode: C**
Anode reaction: **$2I^-$ --------> I_2 + $2e^-$** Cathode reaction: **$2H_2O$ + $2e^-$ -------> $H_2 + 2OH^-$**
Overall reaction: **$2H_2O$ + $2I^-$ -----> H_2 + $2OH^-$ + I_2** MTV: **+0.95 v**

4. The electrorefining of Pb

Anode: **Impure Lead** Cathode: **Pure Lead**
Anode reaction: **Pb-----> Pb^{+2} + $2e^-$** Cathode reaction: **Pb^{2+} + $2e^-$ -----> Pb**

5. Nickel plating an iron nail.

Anode: **Ni** Cathode: **nail**
Anode reaction: **Ni-----> Ni^{+2} + $2e^-$** Cathode reaction: **Ni^{2+} + $2e^-$ -----> Ni**
Possible Electrolyte **$Ni(NO_3)_2$** The -ve side of the power supply is connected to the **nail**

6. Draw an Ag/ Zn electrochemical cell.

Anode: Zn **Cathode: Ag**
Anode reaction: **Zn --------> Zn^{2+} + $2e^-$** Cathode reaction: **Ag^+ + $1e^-$ -------> Ag**

Overall reaction: **Zn + $2Ag^+$ -----> Zn^{2+} + $2Ag$** Voltage: **1.56v**

7. Draw a $KF_{(l)}$ electrolytic cell.

Anode: **C** **Cathode:** **C**

Anode reaction: **$2F^-$ --------> F_2 + $2e^-$** Cathode reaction: **K^+ + e^- -------> K**

Overall reaction: **$2F^-$ + $2K^+$ -----> Cl_2 + K** MTV: **+5.80v**

8. Draw a $KF_{(aq)}$ electrolytic cell.

Anode: **C** **Cathode:** **C**

Anode reaction: **H_2O --------> $2H^+$ + $1/2O_2$ + $2e^-$** Cathode reaction: **$2H_2O$ + $2e^-$ -------> H_2 + $2OH^-$**

Overall reaction: **H_2O -----> H_2 + $1/2O_2$** MTV: **+1.23 v**

9. Draw a $FeI_{2(aq)}$ electrolytic cell.

Anode: **C** **Cathode:** **C**

Anode reaction: **$2I^-$ --------> I_2 + $2e^-$** Cathode reaction: **Fe^{2+} + $2e^-$ -------> Fe**

Overall reaction: **Fe^{2+} + $2I^-$ -----> Fe + I_2** MTV: **+0.99 v**

10. Draw a Cd/Pb electrochemical cell. Cd is not on the reduction chart, however, the Cd electrode gains mass and the total cell potential is .5v. Determine the half-cell potential for Cd.

Anode: **Pb** **Cathode:** **Cd**

Anode reaction: **Pb --------> Pb^{2+} + $2e^-$ 0.13v** Cathode reaction: **Cd^{+2} + $2e^-$ -------> Zn x volts**

Overall reaction: **Pb + Cd^{2+} -----> Pb^{2+} + Cd** Voltage: **0.50v**

0.13 + x = 0.50 **x = 0.37v**

11. Write the overall reaction and describe the anode and cathode for a dry (Leclanche), fuel, alkaline and lead/acid cell.

12. $2HIO_3 + 5H_2SO_3$ ----------> $I_2 + 5H_2SO_4 + H_2O$

oxidizing agent **HIO_3** substance oxidized **H_2SO_3**

substance reduced **HIO_3** reducing agent **H_2SO_3**

13.What is the electrolyte in a fuel cell? **KOH**

14. What is the fuel in a fuel cell? **H_2 and O_2**

15. Describe the **differences** and **similarities** between an electrolytic and electrochemical cell.

Electrolytic	Electrochemical
Uses electricity	**Produces electricity**
Nonspontaneous	**Spontaneous**
Makes chemicals	**Uses chemicals**
Inert carbon electrodes	**Usually has a salt bridge**
The negative electrode is reduction	**The higher metal is reduction**

Cell	anode	anode reaction	cathode	cathode reaction	electrolyte
Leclanche or Common Dry Cell	Zn	Zn-->Zn+2 + 2e-	C	$Mn^{+4} + 1e^-$ -----> Mn^{+3}	NH_4Cl and MnO_2
Alkaline Cell	Zn	Zn-->Zn+2 + 2e-	C	$Mn^{+4} + 1e^-$ -----> Mn^{+3}	KOH and MnO_2
Lead Storage or Car Battery	Pb	Pb ---> Pb^{+2} + $2e^-$	PbO_2	$PbO_2 + HSO_4^- + 3H^+ + 2e^-$ -----> $PbSO_4 + 2H_2O$	H_2SO_4

Oxidation occurs at the anod

e and reduction occurs at the cathode.
Anions migrate to the anode and cations migrate to the cathode.
Electrons go from anode to cathode through the wire.

16. Describe and give two examples of electrowinning. **The electrolysis of water to make H_2 and O_2. The electrolysis of Al_2O_3 to make Al and O_2.**

17. Describe and give one example of electrorefinning. **The electrorefinning of Pb.**

18. List three metals that can be won from aqueous solution. **Pb Au Ag Zn Cu Fe Sn**

19. List three metals that cannot be won from aqueous solution. **Na K Li Ca Mg Al**

20. What is the electrolyte in a fuel cell, alkaline battery, Dry Cell (Leclanche) and lead acid battery?

KOH KOH & MnO_2 NH_4Cl & MnO_2 $PbSO_4$

21. State two metals that can be used to cathodically protect Fe. Describe how they protect iron from corrosion.

Zn and Mg. When attached to Fe they form an electrochemical cell. Zn or Mg is a stronger reducing agent (lower on the chart) and is the anode and Fe is the cathode. Since the cathode is the site of reduction, Fe cannot oxidize or corrode.

22. Write the half reaction that describes the corrosion of iron. **Fe --------> Fe^{2+} $+2e^-$**

23. Write the half reaction that describes the reduction reaction that occurs when iron corrodes in air and water. **$2e^-$ + H_2O + $1/2O_2$ ----------> $2OH^-$**

24. Why does iron corrode faster in salt water? The **salt acts like a salt-bridge and increases the rate of reaction in an electrochemical cell.**

25. Write the anode and cathode reaction in an electrolytic cell with a $CaCl_2$ (l) electrolyte.

Cathode: Ca^{2+} + $2e^-$ ---------> Ca Anode: $2Cl^-$ ----------> Cl_2 + $2e^-$

26. Explain why you would choose Zn or Cu to cathodically protect iron? **Zn. It is a stronger reducing agent than Fe and it will allow Fe to be the cathode, which cannot corrode.**

27. Choose a suitable redox reactant to oxidize Cl^- to ClO_4^- in a redox titration.

MnO_4^- in acid gives a spontaneous reaction as well as a color change from purple to clear.

28. Describe as an electrochemical or electrolytic cell:

a) Fuel cell **electrochemical** b)Charging a car battery **electrolytic**
c) Discharging a car battery **electrochemical** d) Ni plating **electrolytic**
e) Industrial Al production **electrolytic** f) Cl_2 production **electrolytic**

29) Write the anode and cathode reactions.

Cell	anode	anode reaction	cathode	cathode reaction	electrolyte
Cl_2 production	C	$2Cl^-$ ------> Cl_2 + $2e^-$	C	Na^+ + e^- -----> Na	$NaCl_{(l)}$
Leclanche or Common Dry Cell	Zn	Zn-->Zn^{+2} + $2e^-$	C/MnO2	Mn^{+4} +$1e^-$ -----> Mn^{+3}	NH_4Cl and MnO_2
Nickel Plating	Ni	Ni-->Ni^{+2} + $2e^-$	Metal to be plated	Ni^{2+} +$2e^-$ -----> Ni	$Ni(NO_3)_2$
Lead Storage or Car Battery	Pb	Pb ---> Pb^{+2}+ $2e^-$	PbO_2	PbO_2 + SO_4^{-2} + $4OH^{-1}$ + $2e^-$ -----> $PbSO_4$ + $2H_2O$	H_2SO_4
Fuel Cell	C	H_2 + $2OH^-$ ---> $2H_2O$ + $2e^-$	C	O_2 + $2H_2O$ +$4e^-$----> $4OH^-$	KOH

30) Al and AgNO3(aq) are mixed and the surface of the Al darkens. List the two oxidizing agents in decreasing strength. List the two reducing agents in decreasing strength.

Oxidizing Agents **Ag^+** **Al^{3+}**

Reducing Agents **Al** **Ag**