Assign the correct oxidation number to each atom in each of the following. Write the numbers directly above the symbols in each formula, as in the samples.

Samples $\overset{1+}{H_2}\overset{2-}{O}$ (water) $\overset{3-}{N}\overset{1+}{H_4^+}$ (ammonium ion) $\overset{1+}{Na_2}\overset{6+}{S}\overset{2-}{O_4}$ (sodium sulfate)

A. 7. K 8. RbCl 9. Na_2O 10. H_2O_2 11. $MgBr_2$

12. CaS 13. $K_2Cr_2O_7$ 14. H_3PO_4 15. $(NH_4)_3PO_4$ 16. $(NH_4)_2S$

17. $FeCl_2$ 18. $FeCl_3$ 19. FeO 20. N_2 21. N_2O

B. 22. NO 23. N_2O_4 24. NO_2 25. N_2O_3 26. N_2O_5

27. Ag 28. $\cancel{HC_2H_3O_2}$ 29. $CaCO_3$ 30. CO_2 31. CO

32. $Na_2S_2O_3$ 33. S_8 34. Na_2SO_3 35. $MnSO_4$ 36. $KMnO_4$

C. 37. OH^- 38. CO_3^{2-} 39. NO_2^- 40. CrO_4^{2-} 41. $Cr_2O_7^{2-}$

42. IO_3^- 43. IO_4^- 44. $\cancel{SCN^-}$ 45. HSO_4^- 46. PO_4^{3-}

47. HCO_3^- 48. Hg_2^{2+} 49. ClO^- 50. ClO_3^- 51. $S_2O_3^{2-}$

$\boxed{II}$

Determine what gets oxidized & Reduced (16)
Then use table J to determine if the Rxn
is spontaneous or non-spontaneous

4. $I^- + Br_2 \longrightarrow Br^- + I_2$

5. $Zn + CuCl_2 \rightarrow Cu + ZnCl_2$

6. $Cu + AgNO_3 \rightarrow Cu(NO_3)_2 + Ag$

7. $Ag + Au(NO_3)_2 \rightarrow AgNO_3 + Au$

8. $Al + NiCl_2 \longrightarrow Ni + AlCl_3$

9. $Cl_2 + NaF \rightarrow F_2 + NaCl$

10. $F_2 + NaBr \longrightarrow Br_2 + NaF$

Regents Chemistry: Half-Reactions

Name: ______________________________ Date: ________

6

Half-reactions must conserve atoms and charge. That means the total number of atoms of each element must be the same on both sides of the half-reaction and the total charge must be the same on both sides.

For example, we cannot write $O^{-2} \rightarrow O_2$. Neither atoms nor charge is conserved.
First balance the atoms: $2\,O^{-2} \rightarrow O_2$.
Second balance the charge. There is a total of 2(-2) or –4 on the left so there must be negative 4 on the right. $2\,O^{-2} \rightarrow O_2 + 4\,e^-$.

NOTE: OXIDATION NO. OF A PURE element IS ZERO (0)

1) Add electrons to complete the following half-reactions. Identify each half-reaction as oxidation or reduction.

a) Cl_2 + ______ $\rightarrow$ $2\,Cl^{-1}$ ____________

b) Sn^{+4} + ______ $\rightarrow$ Sn^{+2} ____________

c) Pb^{+2} $\rightarrow$ Pb^{+4} + ______ ____________

d) As^{+5} + ______ $\rightarrow$ As^{+2} ____________

e) P_4 $\rightarrow$ $4\,P^{+5}$ + ______ ____________

For the equation:

2) $Mg + S \rightarrow MgS$

a) Assign oxidation numbers for every element in the reactants and products.

b) What is oxidized? ____________

c) What is reduced? ____________

d) Write a half-reaction to represent oxidation:

e) Write a half-reaction to represent reduction:

3) $Zn(s) + 2\,H^{+1}(aq) \rightarrow Zn^{+2}(aq) + H_2(g)$

a) Assign oxidation numbers for every element in the reactants and products.

b) What is oxidized? ____________

c) What is reduced? ____________

d) Write a half-reaction to represent oxidation:

e) WRITE A HALF-REACTION TO REPRESENT REDUCTION

2

~~e) Write a half-reaction to represent reduction:~~

5) $Cl_2(g) + 2\,Br^{-1}(aq) \rightarrow 2\,Cl^{-1}(aq) + Br_2(l)$

a) Assign oxidation numbers for every element in the reactants and products.

b) What is oxidized? ____________________

c) What is reduced? ____________________

d) Write a half-reaction to represent oxidation:

e) Write a half-reaction to represent reduction:

6) $Ca(s) + 2\,H^{+1}(aq) \rightarrow Ca^{+2}(aq) + H_2(g)$

a) Assign oxidation numbers for every element in the reactants and products.

b) What is oxidized? ____________________

c) What is reduced? ____________________

d) Write a half-reaction to represent oxidation:

e) Write a half-reaction to represent reduction:

7) $2\,Al(s) + 3\,Cu^{+2}(aq) \rightarrow 2\,Al^{+3} + 3\,Cu(s)$

a) Assign oxidation numbers for every element in the reactants and products.

b) What is oxidized? ____________________

c) What is reduced? ____________________

d) Write a half-reaction to represent oxidation:

e) Write a half-reaction to represent reduction:

Regents Chemistry: Balancing Redox Reactions

Name: ______________________________ Date: ________

12

In all correctly balanced redox equations, the number of electrons lost by one species is equal to the number of electrons gained by another species. Redox equations are balanced by writing out the oxidation half-reaction and the reduction half-reaction and then balancing the number of electrons lost and gained. After balancing the redox portion of the equation, the remainder can be balanced by inspection.

Balance the following redox equations. <u>SHOW ALL WORK!!!</u>

1	___Sn^0 + ___Ag^{+1} → ___Sn^{+2} + ___Ag^0
2	___Cr^0 + ___Pb^{+2} → ___Cr^{+3} + ___Pb^0
3	___Zn^0 + ___H^{+1} → ___Zn^{+2} + ___H_2
4	___Al^0 + ___Cu^{+2} → ___Al^{+3} + ___Cu^0
5	___Zn^0 + ___Cr^{+3} → ___Zn^{+2} + ___Cr^0
6	___Al^0 + ___O_2 → ___Al_2O_3

1b

Regents Chemistry: Analyzing Skeleton Equations
Name: ______________________________ Date: __________

1. For the skeleton equation:

$$Cr + Cl_2 \rightarrow CrCl_3$$

a) Assign oxidation numbers to all the elements.

b) What is oxidized? ____________________

c) What is reduced? ____________________

d) What is the reducing agent? ____________________

e) What is the oxidizing agent? ____________________

f) Write a half-reaction to represent oxidation:

g) Write a half-reaction to represent reduction:

h) Balance the electrons in the half-reactions and then balance the equation:

2. For the skeleton equation:

$$H_2S + FeCl_3 \rightarrow FeCl_2 + S + HCl$$

a) Assign oxidation numbers to all the elements.

b) What is oxidized? ____________________

c) What is reduced? ____________________

d) What is the reducing agent? ____________________

e) What is the oxidizing agent? ____________________

f) Write a half-reaction to represent oxidation:

g) Write a half-reaction to represent reduction:

h) Balance the electrons in the half-reactions and then balance the equation:

4

In the reaction $Mg + Cl_2 \rightarrow MgCl_2$, the correct half-reaction for the oxidation that occurs is
(1) $Mg + 2e^- \rightarrow Mg^{2+}$ (2) $Cl_2 + 2e^- \rightarrow 2Cl^-$
(3) $Mg \rightarrow Mg^{2+} + 2e^-$ (4) $Cl_2 \rightarrow 2Cl^- + 2e^-$

3. Based on Reference Table N, which of the following ions is most easily oxidized? (1) F^- (2) Cl^- (3) Br^- (4) I^-

4. Based on Reference Table N, which half-cell has a greater reduction potential than the standard hydrogen half-cell?
(1) $Na^+ + e^- \rightarrow Na(s)$ (2) $Ni^{2+} + 2e^- \rightarrow Ni(s)$
(3) $Pb^{2+} + 2e^- \rightarrow Pb(s)$ (4) $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$

5. According to Reference Table N, which will reduce Mg^{2+} to Mg(s)?
(1) Fe(s) (2) Ba(s) (3) Pb(s) (4) Ag(s)

6. Which ion will oxidize Fe?
(1) Zn^{2+} (2) Ca^{2+} (3) Mg^{2+} (4) Cu^{2+}

7. Which ion can be both an oxidizing agent and a reducing agent?
(1) Sn^{2+} (2) Cu^{2+} (3) Al^{3+} (4) Fe^{3+}

8. Which reaction will take place spontaneously?
(1) $Cu + 2H^+ \rightarrow Cu^{2+} + H_2$ (2) $2Au + 6H^+ \rightarrow 2Au^{3+} + 3H_2$
(3) $Pb + 2H^+ \rightarrow Pb^{2+} + H_2$ (4) $2Ag + 2H^+ \rightarrow 2Ag^+ + H_2$

9. Which overall reaction in a chemical cell has the highest net potential (E^0)? -Table N
(1) $Zn(s) + 2H^+ \rightarrow Zn^{2+} + H_2(g)$ (2) $Ni(s) + 2H^+ \rightarrow Ni^{2+} + H_2(g)$
(3) $Mg(s) + 2H^+ \rightarrow Mg^{2+} + H_2(g)$ (4) $Sn(s) + 2H^+ \rightarrow Sn^{2+} + H_2(g)$

10. Given the reaction: Table N

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

The potential for (E^0) for the overall reaction is (1) 1.53 V (2) 1.79 V (3) 2.93 V (4) 3.71 V

Base your answers to questions 11 and 12 on the following reaction:

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

11. Which species undergoes a loss of electrons?
(1) Mg(s) (2) $Ag^+(aq)$ (3) $Mg^{2+}(aq)$ (4) Ag(s)

2. What is the cell voltage (E^0) for the overall reaction? Table N
(1) + 1.57 V (2) +2.37 V (3) +3.17 V (4) +3.97 V

1. Given the redox reaction: $Ni + Sn^{4+} \rightarrow Ni^{2+} + Sn^{2+}$
Which species has been oxidized? (1) Ni (2) Sn^{4+} (3) Ni^{2+} (4) Sn^{2+}

2. In which substance is the oxidation number of nitrogen zero?
(1) N_2 (2) NH_3 (3) NO_2 (4) N_2O

3. What is the oxidation number of Pt in K_2PtCl_6? (1) −2 (2) +2 (3) −4 (4) +4

4. In the reaction $2H_2S + 3O_2 \rightarrow 2SO_2 + 2H_2O$, the oxidizing agent is (1) oxygen (2) water (3) sulfur dioxide (4) hydrogen sulfide

5. In the reaction $2Na + 2H_2O \rightarrow 2Na^+ + 2OH^- + H_2$, the substance oxidized is (1) H_2 (2) H^+ (3) Na (4) Na^+

6. Given the reaction: $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$
The reducing agent is (1) Cu^0 (2) N^{+5} (3) Cu^{+2} (4) N^{+2}

7. Given the reaction: $Sn^{2+}(aq) + 2Fe^{3+}(aq) \rightarrow Sn^{4+}(aq) + 2Fe^{2+}(aq)$
The oxidizing agent in this reaction is (1) Sn^{2+} (2) Fe^{3+} (3) Sn^{4+} (4) Fe^{2+}

In 1–19, refer to Reference Table N on page 110.

1. Given the reaction: Table N

$$2Cr(s) + 3Cu^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Cu(s)$$

The potential difference (E^0) of the cell is (1) 0.40 v (2) 1.08 v (3) 1.25 v (4) 2.50 v

2. Given the reaction: Table N

$$2Na(s) + Cl_2(g) \rightarrow 2Na^+ + 2Cl^-$$

What is the potential (E^0) for the overall reaction? (1) −1.35 v (2) +1.35 v (3) −4.07 v (4) +4.07 v

3. Given the reaction: Table N

$$2Au^{3+}(aq) + 3Ni^0 \rightarrow 2Au^0 + 3Ni^{2+}(aq)$$

The cell potential (E^0) for the overall reaction is (1) 3.75 v (2) 2.25 v (3) 1.76 v (4) 1.25 v

4. What is the potential (E^0) for the reaction Table N

$$Mg(s) + Br_2(\ell) \rightarrow Mg^{2+}(aq) + 2Br^-(aq)$$

(1) 1.06 v (2) 1.31 v (3) 2.37 v (4) 3.46 v

5. Which species can act either as an oxidizing agent or a reducing agent? (1) Na^0 (2) Fe^{2+} (3) Sn^0 (4) Zn^{2+}

6. Given the reaction:

$$Zn(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Zn^{2+}(aq) + 2Cl^-(aq) + H_2(g)$$

Which species is oxidized?
(1) Zn(s) (2) H^+ (aq) (3) Cl^- (aq) (4) $H_2(g)$

7. Which pair will react spontaneously at 298 K? (1) $Cl_2 + F^-$ (2) $I_2 + Br^-$ (3) $F_2 + I^-$ (4) $Br_2 + Cl^-$

8. Which pair will react spontaneously at 298 K? (1) $Cu + H_2O$ (2) $Ag + H_2O$ (3) $Ca + H_2O$ (4) $Au + H_2O$

9. What is the standard reduction potential for the $Cu^{2+}(aq)/Cu$ half-cell? (1) +0.52 v (2) +0.34 v (3) −0.52 v (4) −0.34 v

10. Which species can oxidize Sn^{2+} to Sn^{4+}? (1) Ag^+ (2) Fe^0 (3) Al^{3+} (4) H_2O

11. Which of the following Group 17 elements is the strongest oxidizing agent? (1) I_2 (2) Br_2 (3) Cl_2 (4) F_2

12. Which of the following alkaline earth elements is the strongest reducing agent? (1) Mg (2) Sr (3) Ca (4) Ba

13. Which species can be reduced by Zn? (1) Na^+ (2) H^+ (3) Ca^{2+} (4) Mg^{2+}

14. What is the E^0 for the chemical cell whose net reaction is

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

(1) 0.66 volt (2) 0.79 volt (3) 0.94 volt (4) 1.09 volt

15. Given the reaction:

$$\underline{\ ?\ } + Ni^{2+}(aq) \rightarrow \underline{\ ?\ } + Ni$$

If this reaction is spontaneous, the missing reactant could be (1) Zn^0 (2) Pb^0 (3) Cu^0 (4) Sn^0

10. In the decomposition reaction $2LiH \rightarrow 2Li + H_2$ (1) Li^+ is oxidized. (2) Li^0 is oxidized. (3) $H_2{}^0$ is reduced. (4) H^- is oxidized.

11. In the double replacement reaction

$$AgNO_3 + NaCl \rightarrow AgCl(s) + NaNO_3$$

(1) $AgNO_3$ is reduced. (2) $AgNO_3$ is oxidized. (3) NaCl is reduced. (4) $AgNO_3$ is neither reduced nor oxidized.