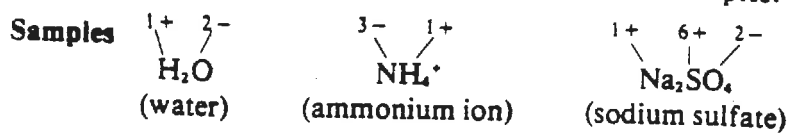


Assign the correct oxidation number to each atom in each of the following. Write the numbers directly above the symbols in each formula, as in the samples.



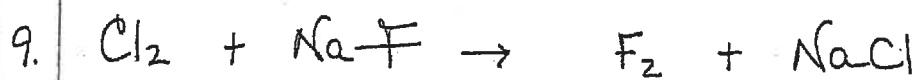
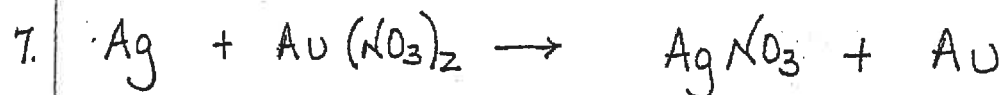
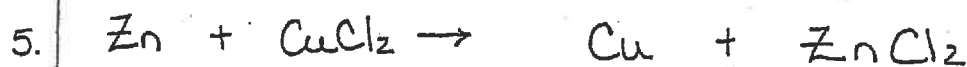
- A.
- |                       |   |                                    |   |                                       |
|-----------------------|---|------------------------------------|---|---------------------------------------|
| 7. K                  | 8. RbCl   | 9. Na <sub>2</sub> O               | 10. H <sub>2</sub> O <sub>2</sub>                   | 11. MgBr <sub>2</sub>                 |
| 12. CaS               | 13. K <sub>2</sub> Cr <sub>2</sub> O <sub>7</sub> | 14. H <sub>3</sub> PO <sub>4</sub> | 15. (NH <sub>4</sub> ) <sub>3</sub> PO <sub>4</sub> | 16. (NH <sub>4</sub> ) <sub>2</sub> S |
| 17. FeCl <sub>2</sub> | 18. FeCl <sub>3</sub>                             | 19. FeO                            | 20. N <sub>2</sub>                                  | 21. N <sub>2</sub> O                  |
- B.
- |   |                                    |                                     |                                   |                                   |
|---|------------------------------------|-------------------------------------|-----------------------------------|-----------------------------------|
| 22. NO  | 23. N <sub>2</sub> O <sub>4</sub>  | 24. NO <sub>2</sub>                 | 25. N <sub>2</sub> O <sub>3</sub> | 26. N <sub>2</sub> O <sub>5</sub> |
| 27. Ag  | 28. <del>HCl, H<sub>2</sub>O</del> | 29. CaCO <sub>3</sub>               | 30. CO <sub>2</sub>               | 31. CO                            |
| 32. Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> | 33. S <sub>8</sub>                 | 34. Na <sub>2</sub> SO <sub>3</sub> | 35. MnSO <sub>4</sub>             | 36. KMnO <sub>4</sub>             |
- C.
- |                                   |                                   |                                  |                                    |  |
|-----------------------------------|-----------------------------------|----------------------------------|------------------------------------|--|
| 37. OH <sup>-</sup>               | 38. CO <sub>3</sub> <sup>2-</sup> | 39. NO <sub>2</sub> <sup>-</sup> | 40. CrO <sub>4</sub> <sup>2-</sup> | 41. Cr <sub>2</sub> O <sub>7</sub> <sup>2-</sup> |
| 42. IO <sub>3</sub> <sup>-</sup>  | 43. IO <sub>4</sub> <sup>-</sup>  | 44. <del>SCN<sup>-</sup></del>   | 45. HSO <sub>4</sub> <sup>-</sup>  | 46. PO <sub>4</sub> <sup>3-</sup>                |
| 47. HCO <sub>3</sub> <sup>-</sup> | 48. Hg <sub>2</sub> <sup>2+</sup> | 49. ClO <sup>-</sup>             | 50. ClO <sub>3</sub> <sup>-</sup>  | 51. S <sub>2</sub> O <sub>3</sub> <sup>2-</sup>  |

*w*

11

Determine what gets oxidized & reduced  
Then use table J to determine if the rxn  
is spontaneous or non-spontaneous

16



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)

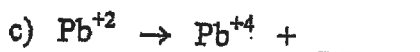
1) Add electrons to complete the following half-reactions. Identify each half-reaction as oxidation or reduction.



\_\_\_\_\_



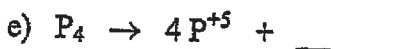
\_\_\_\_\_



\_\_\_\_\_

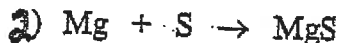


\_\_\_\_\_



\_\_\_\_\_

2) For the equation:



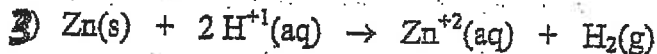
a) Assign oxidation numbers for every element in the reactants and products.

b) What is oxidized? \_\_\_\_\_

c) What is reduced? \_\_\_\_\_

d) Write a half-reaction to represent oxidation:

e) Write a half-reaction to represent reduction:



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

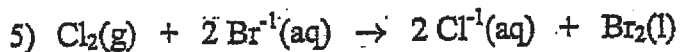
b) What is oxidized? \_\_\_\_\_

c) What is reduced? \_\_\_\_\_

d) Write a half-reaction to represent oxidation:

e) WRITE A HALF-REACTION TO REPRESENT REDUCTION

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



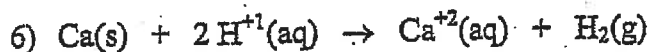
a) Assign oxidation numbers for every element in the reactants and products.

b) What is oxidized? \_\_\_\_\_

c) What is reduced? \_\_\_\_\_

d) Write a half-reaction to represent oxidation:

e) Write a half-reaction to represent reduction:



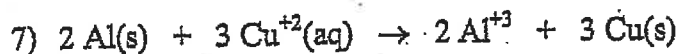
a) Assign oxidation numbers for every element in the reactants and products.

b) What is oxidized? \_\_\_\_\_

c) What is reduced? \_\_\_\_\_

d) Write a half-reaction to represent oxidation:

e) Write a half-reaction to represent reduction:



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

b) What is oxidized? \_\_\_\_\_

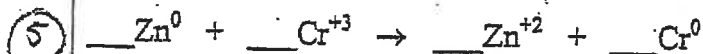
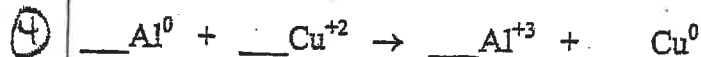
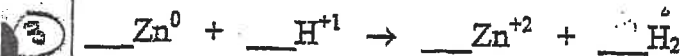
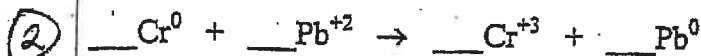
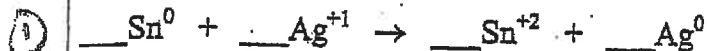
c) What is reduced? \_\_\_\_\_

d) Write a half-reaction to represent oxidation:

e) Write a half-reaction to represent reduction:

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. SHOW ALL WORK!!!



Name: \_\_\_\_\_ Date: \_\_\_\_\_

1. For the skeleton equation:



- 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:



- 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:

1. In the reaction  $\text{Mg} + \text{Cl}_2 \rightarrow \text{MgCl}_2$ , the correct half-reaction for the oxidation that occurs is

- (1)  $\text{Mg} + 2\text{e}^- \rightarrow \text{Mg}^{2+}$  (2)  $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$   
(3)  $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$  (4)  $\text{Cl}_2 \rightarrow 2\text{Cl}^- + 2\text{e}^-$

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

4. Based on Reference Table N, which half-cell has a greater reduction potential than the standard hydrogen half-cell?

- (1)  $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}(\text{s})$  (2)  $\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}(\text{s})$   
(3)  $\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}(\text{s})$  (4)  $\text{Sn}^{4+} + 2\text{e}^- \rightarrow \text{Sn}^{2+}$

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

6. Which ion will oxidize  $\text{Fe}$ ?

- (1)  $\text{Zn}^{2+}$  (2)  $\text{Ca}^{2+}$  (3)  $\text{Mg}^{2+}$  (4)  $\text{Cu}^{2+}$

7. Which ion can be both an oxidizing agent and a reducing agent?

- (1)  $\text{Sn}^{2+}$  (2)  $\text{Cu}^{2+}$  (3)  $\text{Al}^{3+}$  (4)  $\text{Fe}^{3+}$

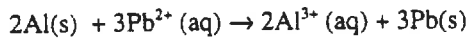
8. Which reaction will take place spontaneously?

- (1)  $\text{Cu} + 2\text{H}^+ \rightarrow \text{Cu}^{2+} + \text{H}_2$  (2)  $2\text{Au} + 6\text{H}^+ \rightarrow 2\text{Au}^{3+} + 3\text{H}_2$   
(3)  $\text{Pb} + 2\text{H}^+ \rightarrow \text{Pb}^{2+} + \text{H}_2$  (4)  $2\text{Ag} + 2\text{H}^+ \rightarrow 2\text{Ag}^+ + \text{H}_2$

9. Which overall reaction in a chemical cell has the highest net potential ( $E^\circ$ )? -Table N

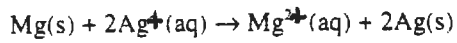
- (1)  $\text{Zn}(\text{s}) + 2\text{H}^+ \rightarrow \text{Zn}^{2+} + \text{H}_2(\text{g})$  (2)  $\text{Ni}(\text{s}) + 2\text{H}^+ \rightarrow \text{Ni}^{2+} + \text{H}_2(\text{g})$   
(3)  $\text{Mg}(\text{s}) + 2\text{H}^+ \rightarrow \text{Mg}^{2+} + \text{H}_2(\text{g})$  (4)  $\text{Sn}(\text{s}) + 2\text{H}^+ \rightarrow \text{Sn}^{2+} + \text{H}_2(\text{g})$

10. Given the reaction: Table N



The potential for ( $E^\circ$ ) 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:



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

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

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

12. In which substance is the oxidation number of nitrogen zero?  
(1)  $\text{N}_2$  (2)  $\text{NH}_3$  (3)  $\text{NO}_2$  (4)  $\text{N}_2\text{O}$

13. What is the oxidation number of Pt in  $\text{K}_2\text{PtCl}_6$ ? (1) -2 (2) +2 (3) -4 (4) +4

14. In the reaction  $2\text{H}_2\text{S} + 3\text{O}_2 \rightarrow 2\text{SO}_2 + 2\text{H}_2\text{O}$ , the oxidizing agent is (1) oxygen (2) water (3) sulfur dioxide (4) hydrogen sulfide

15. In the reaction  $2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{Na}^+ + 2\text{OH}^- + \text{H}_2$ , the substance oxidized is (1)  $\text{H}_2$  (2)  $\text{H}^+$  (3)  $\text{Na}^+$  (4)  $\text{Na}$

16. Given the reaction:  $3\text{Cu} + 8\text{HNO}_3 \rightarrow 3\text{Cu}(\text{NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}$   
The reducing agent is (1)  $\text{Cu}^0$  (2)  $\text{N}^{+5}$  (3)  $\text{Cu}^{+2}$  (4)  $\text{N}^{+2}$

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

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

1. Given the reaction: Table N  
 $2\text{Cr}(\text{s}) + 3\text{Cu}^{2+}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 3\text{Cu}(\text{s})$

The potential difference ( $E^\circ$ ) 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  
 $2\text{Na}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{Na}^+(\text{aq}) + 2\text{Cl}^-(\text{aq})$

What is the potential ( $E^\circ$ ) 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  
 $2\text{Au}^{3+}(\text{aq}) + 3\text{Ni}^0 \rightarrow 2\text{Au}^0 + 3\text{Ni}^{2+}(\text{aq})$

The cell potential ( $E^\circ$ ) 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 ( $E^\circ$ ) for the reaction Table N  
 $\text{Mg}(\text{s}) + \text{Br}_2(\ell) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{Br}^-(\text{aq})$

(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)  $\text{Na}^0$  (2)  $\text{Fe}^{2+}$  (3)  $\text{Sn}^0$  (4)  $\text{Zn}^{2+}$

6. Given the reaction:  
 $\text{Zn}(\text{s}) + 2\text{H}^+(\text{aq}) + 2\text{Cl}^-(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) + \text{H}_2(\text{g})$

Which species is oxidized?

- (1)  $\text{Zn}(\text{s})$  (2)  $\text{H}^+(\text{aq})$  (3)  $\text{Cl}^-(\text{aq})$  (4)  $\text{H}_2(\text{g})$

7. Which pair will react spontaneously at 298 K? (1)  $\text{Cl}_2 + \text{F}^-$  (2)  $\text{I}_2 + \text{Br}^-$  (3)  $\text{F}_2 + \text{I}^-$  (4)  $\text{Br}_2 + \text{Cl}^-$

8. Which pair will react spontaneously at 298 K? (1)  $\text{Cu} + \text{H}_2\text{O}$  (2)  $\text{Ag} + \text{H}_2\text{O}$  (3)  $\text{Ca} + \text{H}_2\text{O}$  (4)  $\text{Au} + \text{H}_2\text{O}$

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

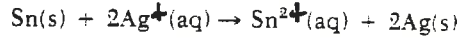
10. Which species can oxidize  $\text{Sn}^{2+}$  to  $\text{Sn}^{4+}$ ? (1)  $\text{Ag}^+$  (2)  $\text{Fe}^0$  (3)  $\text{Al}^{3+}$  (4)  $\text{H}_2\text{O}$

11. Which of the following Group 17 elements is the strongest oxidizing agent? (1)  $\text{I}_2$  (2)  $\text{Br}_2$  (3)  $\text{Cl}_2$  (4)  $\text{F}_2$

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

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

14. What is the  $E^\circ$  for the chemical cell whose net reaction is



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

15. Given the reaction:  
 $\underline{\quad} + \text{Ni}^{2+}(\text{aq}) \rightarrow \underline{\quad} + \text{Ni}$

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

10. In the decomposition reaction  $2\text{LiH} \rightarrow 2\text{Li} + \text{H}_2$  (1)  $\text{Li}^+$  is oxidized. (2)  $\text{Li}^0$  is oxidized. (3)  $\text{H}_2^0$  is reduced. (4)  $\text{H}^+$  is oxidized.

11. In the double replacement reaction  
 $\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3$

(1)  $\text{AgNO}_3$  is reduced. (2)  $\text{AgNO}_3$  is oxidized. (3)  $\text{NaCl}$  is reduced. (4)  $\text{AgNO}_3$  is neither reduced nor oxidized.