Oxidation-Reduction and Electrochemistry Exam Review

Assign each element in the following compounds its correct oxidation #.

1. HNO_3

   +1 +5 −2

   $1 + x + 3(-2) = 0$
   $1 + x - 6 = 0$
   $x - 5 = 0$
   $x = +5$

2. SO_4^{2-}

   +6 −2

   $x + 4(-2) = -2$
   $x - 8 = -2$
   $x = +6$

3. H_2CO_3

   +1 +4 −2

   ~~2x+~~
   $2(1) + x + 3(-2) = 0$
   $2 + x - 6 = 0$
   $x - 4 = 0$
   $x = +4$

4. H_2SeO_3

   +1 +4 −2

   $2(1) + x + 3(-2) = 0$
   $2 + x - 6 = 0$
   $x - 4 = 0$
   $x = +4$

5. Zn

   0

6. $H_2\,(g)$

   0

7. $Na_2Cr_2O_7$

   +1 +6 −2

   $2(1) + 2x + 7(-2) = 0$
   $2 + 2x - 14 = 0$
   $2x - 12 = 0$
   $\frac{2x}{2} = \frac{12}{2}$
   $x = +6$

8. Al_2O_3

   +3 −2

   $2x + 3(-2) = 0$
   $2x - 6 = 0$
   $2x = 6$
   $x = 3$

9. $SbCl_5$

   +5 −1

   $x + 5(-1) = 0$
   $x - 5 = 0$
   $x = 5$

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Assign Oxidation #'s to each element in the following equations. Identify the element that is oxidized, the element that is reduced, the oxidizing agent and the reducing agent.

Charge ↑: lost e^-

Charge ↓: gain e^-

$x - 8 = -1$

$x = +7$

$$\overset{+2}{Fe^{2+}} + \overset{+7\,-2}{MnO_4^-} \rightarrow \overset{+2}{Mn^{2+}} + \overset{+3}{Fe^{3+}}$$

10. $Fe^{2+} + MnO_4^- \rightarrow Mn^{2+} + Fe^{3+}$

Fe: $+2 \rightarrow +3$, lost e^-, ox, RA: Fe^{2+}

Mn: $+7 \rightarrow +2$, gain e^-, red, OA: MnO_4^-

Oxidized	Fe	Reduced	Mn
Oxidizing agent	MnO_4^-	Reducing Agent	Fe^{2+}

$$\overset{+2\,-2}{CuS} + \overset{+5\,-2}{NO_3^-} \rightarrow \overset{+2}{Cu^{2+}} + \overset{+4\,-2}{NO_2} + \overset{0}{S}$$

11. $CuS + NO_3^- \rightarrow Cu^{2+} + NO_2 + S$

$x + 3(-2) = -1$

$x - 6 = -1$

$x = +5$

$x + 2(-2) = 0$

$x - 4 = 0$

$x = +4$

S: $-2 \rightarrow 0$, lost e^-, ox, RA: CuS

N: $+5 \rightarrow +4$, gain e^-, red, OA: NO_3^-

Oxidized	S	Reduced	N
Oxidizing agent	NO_3^-	Reducing Agent	CuS

+5 -2 | +4 -2 | -1 | +7 -2

12. $\boxed{BrO_3^- + MnO_2} \rightarrow Br^- + MnO_4^-$

$x + 3(-2) = -1$
$x - 6 = -1$
$x = +5$

$x + 4(-2) = -1$
$x - 8 = -1$
$x = +7$

Br: +5 → -1, gain e^-, red, OA: BrO_3^-
Mn: +4 → +7, lost e^-, ox, RA: MnO_2

Oxidized	Mn	Reduced	Br
Oxidizing agent	BrO_3^-	Reducing Agent	MnO_2

+2 +6 -2 | +1 -2 | 0 | +4 -2 | +1 +6 -2

13. $2CdSO_4 + H_2O \rightarrow Cd + CdO_2 + 2H_2SO_4$

$2(1) + x + 4(-2) = 0$
$2 + x - 8 = 0$
$x - 6 = 0$
$x = 6$

$x + 6 + 4(-2) = 0$
$x + 6 - 8 = 0$
$x - 2 = 0$
$x = +2$

Cd: +2 → 0, gain e^-, red
Cd: +2 → +4, lost e^-, ox

+6 -2 SO_4^{2-}
+5 -2 NO_3^-
+5 -2 PO_4^{3-}

Oxidized	Cd	Reduced	Cd
Oxidizing agent	$CdSO_4$	Reducing Agent	$CdSO_4$

14. $\boxed{\overset{-4\ +1}{CH_4} + \overset{0}{O_2} \rightarrow}$ $\overset{+1\ -2}{H_2O} + \overset{+4\ -2}{CO_2}$

C: -4 → +4, lost e⁻, ox, RA: CH4

O: 0 → -2, gain e⁻, red, OA: O2

Oxidized	C	Reduced	O
Oxidizing agent	O2	Reducing Agent	CH4

15. $\boxed{\overset{0}{Cu} + 2\overset{+1\ x\ -2}{Ag\underline{NO_3}} \rightarrow}$ $\overset{+2\ x\ -2}{Cu(\underline{NO_3})_2} + 2\overset{0}{Ag}$

NO_3^{-} (O: -2)

$x + 3(-2) = -1$

$x - 6 = -1$

$x = +5$

~~NO3~~

$x + 5 - 6 = 0$

$x - 1 = 0$

$x = 1$

$x + 2(5) + 6(-2) = 0$

$x + 10 - 12 = 0$

$x - 2 = 0$ $x = +2$

Oxidized	Cu	Reduced	Ag
Oxidizing agent	AgNO3	Reducing Agent	Cu

Cu: 0 → +2, lost e⁻, ox, RA: Cu

Ag: +1 → 0, gain e⁻, red, OA: AgNO3

Fill in the blanks: Oxidation reduction reactions involve the movement of electrons from one atom to another. The substance that loses electrons is oxidized and the substance that gains electrons is reduced. When an atom loses electrons, it takes on a positive charge. Some elements lose electrons more easily than others. Alkali metals metals tend to easily lose electrons and take on a +1 charge. Jewelry metals are not very reactive because they do not tend to lose electrons. Halogens are more easily reduced than oxidized. That means that they gain electrons. An element's oxidation number describes how many electrons it has lost or gained. When an element is in its standard state, its oxidation number is zero. During any redox reaction, one element loses electrons and another gains electrons. These half-reactions can be linked together to form a voltaic/galvanic cell, or battery. The half reactions are connected by a salt bridge, or solution of ions that regulates charge in the battery. The electrons in a voltaic cell flow from the anode to the cathode. The cell potential/charge of the battery can be found by adding up the potential values of each half cell. The charge of the battery is measured in units of volts.

16. A student wants to design a battery with copper and zinc metals and metal solutions.

+0.34 Cu Cu^{2+} Zn Zn^{2+} -0.76

the most neg. cell potential will be oxidized!

a. Which of these two metals is most easily oxidized? Zn

b. Which of these two metals is most easily reduced? Cu

c. Using your table of reduction potentials in your notes, write the oxidation and reduction half reactions below along with their voltage potentials. (Don't forget to flip the oxidation half reaction. Also, use the copper reaction that involved two electrons.)

make the ox. positive!

- Oxidation reaction $Zn \rightarrow Zn^{2+} + 2e^-$ $E^0 =$ +0.76V
- Reduction reaction $Cu^{2+} + 2e^- \rightarrow Cu$ $E^0 =$ 0.34V

add!

d. What is the total charge of the cell? 1.10V

e. Draw a salt bridge and wire to complete the diagram below. Label the anode and the cathode. Indicate the direction the electrons would flow in the wire with arrows.

anode/cathode: An Ox
Red Cat

e⁻ flow: FAT CAT

Zn: ox
Cu: red