1. Determine the oxidation number of the elements in each of the following compounds:   
   (See the rules for assigning oxidation numbers on the back of this page)

a. H2CO3  H= C= O=  
  
b. N2  N=  
  
c. Zn(OH)42-  Zn= O= H=  
  
d. NO2-  N= O=  
  
e. LiH  Li= H=   
  
f. Fe3O4  Fe= O=

1. Identify the species being oxidized and reduced in each of the following reactions: (Look at oxidation numbers.)  
     
    Oxidized Reduced

a. 2 Cr+ + Sn4+ http://www.chemistry.wustl.edu/~coursedev/Online%20tutorials/arrow.gif Cr3+ + Sn2+   
  
b. 3 Hg2+ + 2 Fe (s) http://www.chemistry.wustl.edu/~coursedev/Online%20tutorials/arrow.gif 3 Hg2 + 2 Fe3+   
  
c. 2 As (s) + 3 Cl2 (g) http://www.chemistry.wustl.edu/~coursedev/Online%20tutorials/arrow.gif 2 AsCl3

1. Would you use an oxidizing agent or reducing agent in order for the following reactions to occur? (Look at oxidation numbers. An oxidizing agent (or oxidant) is one that causes oxidation. It is reduced in the reaction. A reducing agent (reductant) is one that causes the reduction in the redox reaction. It is oxidized in the reaction.)

a. ClO3- http://www.chemistry.wustl.edu/~coursedev/Online%20tutorials/arrow.gif ClO2    
  
b. SO42- http://www.chemistry.wustl.edu/~coursedev/Online%20tutorials/arrow.gif S2-    
  
c. Mn2+ http://www.chemistry.wustl.edu/~coursedev/Online%20tutorials/arrow.gif MnO2    
  
d. Zn http://www.chemistry.wustl.edu/~coursedev/Online%20tutorials/arrow.gif ZnCl2 

1. Write balance equations for the following redox reactions: (Again start by looking at the oxidation numbers. Then work out the half reactions. Balance these equations and combine. )

Oxidation half reaction Reduction half reaction

a. \_\_\_\_NaBr + \_\_\_\_\_Cl2 http://www.chemistry.wustl.edu/~coursedev/Online%20tutorials/arrow.gif \_\_\_\_\_NaCl + \_\_\_\_\_Br2

b. \_\_\_\_\_Fe2O3 + \_\_\_\_\_CO http://www.chemistry.wustl.edu/~coursedev/Online%20tutorials/arrow.gif \_\_\_\_\_Fe + \_\_\_\_\_CO2

c. \_\_\_\_\_CO + \_\_\_\_\_I2O5 http://www.chemistry.wustl.edu/~coursedev/Online%20tutorials/arrow.gif \_\_\_\_\_CO2 + \_\_\_\_\_I2 

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|  | |  | 1. The oxidation number for an atom in its elemental form is always zero.    * A substance is elemental if both of the following are true:      + only one kind of atom is present      + charge = 0    * Examples:      + S8: The oxidation number of S = 0      + Fe: The oxidation number of Fe = 0 2. The oxidation number of a monoatomic ion = charge of the monatomic ion.    * Examples:      + Oxidation number of S2- is -2.      + Oxidation number of Al3+ is +3. 3. The oxidation number of all Group 1A metals = +1 (unless elemental). 4. The oxidation number of all Group 2A metals = +2 (unless elemental). 5. Hydrogen (H) has two possible oxidation numbers:    * +1 when bonded to a nonmetal    * -1 when bonded to a metal 6. Oxygen (O) has two possilbe oxidation numbers:    * -1 in peroxides (O22-)....pretty uncommon    * -2 in all other compounds...most common 7. The oxidation number of fluorine (F) is always -1. 8. The sum of the oxidation numbers of all atoms (or ions) in a neutral compound = 0. 9. The sum of the oxidation numbers of all atoms in a polyatomic ion = charge on the polyatomic ion.   LEO the lion says GER  Losing Electrons is Oxidation (LEO). Gaining Electrons is Reduction (GER).  *To determine if a redox reaction has occurred and to identify the element oxidized and the element reduced:*   1. Assign oxidation numbers to all atoms in the equation.(NOTE: Ignore the coefficients in the equation. They are not important when determining the oxidation numbers.) 2. Compare oxidation numbers from the reactant side to the product side of the equation.    * If a redox reaction has occurred, you will find that the oxidation numbers of two (no more/no less) elements have changed from the reactant side to the product side. 3. The element oxidized is the one whose oxidation number increased (lost electrons) 4. The element reduced is the one whose oxidation number decreased (gained electrons)   ***Example***: Determine if a redox reaction has occurred. If so, identify the element oxidized and the element reduced.  I2O5 (s) + 5 CO (g) --> I2 (s) + 5 CO2 (g)  ***Solution***: Assign oxidation numbers to each element:   |  |  |  |  |  |  |  | | --- | --- | --- | --- | --- | --- | --- | | I2O5 (s) | + | 5 CO (g) | --> | I2 (s) | + | 5 CO2 (g) |  |  |  |  |  |  |  |  |  |  |  |  |  | | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | | +5 | -2 |  | +2 | -2 |  | 0 |  |  |  | +4 | -2 |   Compare oxidation numbers from side to side:   * The oxidation number of I decreased from +5 in I2O5 to 0 in I2. * The oxidation number of C increased from +2 in CO to +4 in CO2.   + Since the oxidation numbers of two elements changed from side to side, a redox reaction DID occur.   Since the oxidation number of I decreased from +5 to 0 by gaining 5 electrons, I was reduced (GER). Since the oxidation number of C increased from +2 to +4 by losing 2 electrons, C was oxidized (LEO). |