

## I. REVIEW OF OXIDATION AND REDUCTION IN ACID OR BASE CONDITIONS

**Oxidation State** (also known as *oxidation number* or *valence state*), is a quick way of assessing the oxidation or reduction of particular atoms according to a prescribed set of rules. [*Oxidation state* is not the same as *formal charge*, which tries to estimate the real charge distribution of a molecule or ion among its constituent atoms. Formal charge uses information about single bonds, multiple bonds and octet and non-octet structures. *Dipole moment*, the actual distribution of charges in a molecule or ion, must be experimentally measured.]

**Empirical Method of Determining Oxidation State** ( These rules must be memorized and applied **in order**, the first rule has the highest priority, the last has the least priority.)

- 1) Elements in elemental form are 0 (e.g. Fe(s), O<sub>2</sub>(g), etc.)
- 2) The sum of the oxidation states = the overall charge of the molecule/polyatomic ion. (e.g. ClO<sub>4</sub><sup>-</sup>, Cl is +7, O is -2, 7+(4x-2) = -1)
- 3) All group I metals **and** Ag are +1
- 4) F is -1
- 5) All group II metals **and** Zn, Cd are +2
- 6) H is +1
- 7) O is -2

### Some Definitions

**Reduction:** The charge (oxidation state) is reduced by gaining electrons. (e.g., in Fe<sup>+2</sup> + 2e<sup>-</sup> → Fe(s), the charge goes from +2 to 0 and so the Fe<sup>+2</sup> is reduced to Fe(s).)

**Oxidation:** The charge (oxidation state) increases by losing electrons. (eg., Ag(s) → Ag<sup>+</sup> + 1e<sup>-</sup>)

**Reducing agent:** The reducing agent is the compound which contains the element that is oxidized. When reducing another element or compound, the reducing agent must give electrons away so it becomes oxidized.

**Oxidizing agent:** The oxidizing agent is the compound which contains the element that is reduced.

**Redox Reaction:** A balanced chemical reaction consisting of both an oxidation and a reduction. (i.e., the sum of an oxidation 1/2 reaction and a reduction 1/2 reaction.) These are electron transfer reactions.

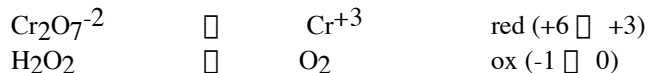
### Balancing Redox Reactions

A six step method:  
**HEIOHEN**

H	EI	O	H	E	N
<b>Half rxn's:</b>	<b>Elements:</b>	<b>Oxygen:</b>	<b>Hydrogen:</b>	<b>Electrons:</b>	<b>Number of electrons:</b>
break up into	balance the Elements	balance O	balance H	add the e <sup>-</sup> to	multiply so the Number
Half (1/2) rxn's	except O and H	with H <sub>2</sub> O	with H <sup>+</sup>	balance charge	of e <sup>-</sup> is the same in both

**Example:** Balance the following reaction in acid: Cr<sub>2</sub>O<sub>7</sub><sup>-2</sup> + H<sub>2</sub>O<sub>2</sub> → Cr<sup>+3</sup> + O<sub>2</sub>

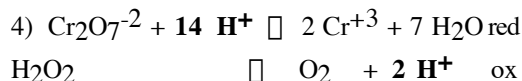
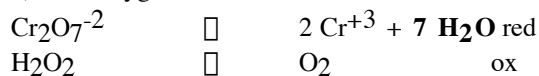
1) **H:** 1/2 rxn's



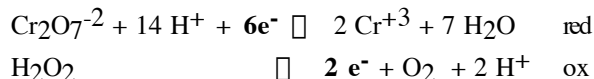
2) **EI:** elements



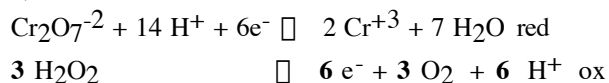
3) **O:** oxygen



5) **E:** electrons Be sure to use the **net** (total) charge on each sides. Multiply coefficients by charges to get the net.



6) **N:** number



7) Now add them up.



**8) Always check that in your answer all electrons cancel and atoms and charges balance!** Also check to see if you can divide by a common divisor (see if you have the lowest possible coefficients) and if you can collect terms.

### Balancing in Base

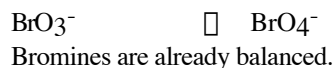
The procedure is the same as with acid except that after you add the  $\text{H}^+$  to balance the hydrogens, add the **same number of  $\text{OH}^-$  to both sides and form water on one side.**

**Example:** Balance the following 1/2 rxn in base:  $\text{BrO}_3^- \square \text{BrO}_4^-$

1) **H:** (The 1/2 rxn is given in this case)



2) **El:** elements



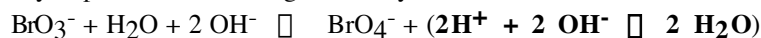
3) **O:** oxygen



4) **H:** hydrogen (but because its really in base, after adding  $\text{H}^+$ , then add  $\text{OH}^-$  to **both** sides.)



If you put  $\text{H}^+$  and  $\text{OH}^-$  together, they will form water.



Cancel the waters and continue.



5) **E:** electrons

