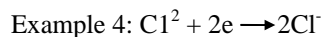
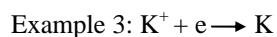
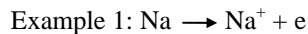


OXIDATION REDUCTION

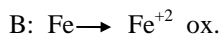
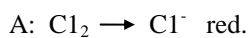
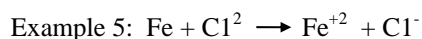
Section I



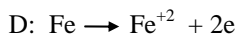
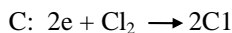
1. The oxidation number of group I A is always (+) 1.
2. The oxidation number of group II A is always (+) 2.
3. The oxidation number of oxygen is -2 exception peroxides. In peroxides the oxidation number of oxygen is -1. Example of peroxides: H_2O_2 , Na_2O_2 .
4. The oxidation number of hydrogen is +1 except in hydrides. In hydrides the oxidation number of H is -1. Example of hydrides: NaH , LiH .
5. The algebraic sum of the oxidation number of all atoms in a neutral compound is zero.
6. The oxidation number of elements in its free state is zero.
7. The oxidation number of the halogens is -1 when they combine with hydrogen or with a metal in a binary compound.

We will discuss the ion-electron method first. The following steps are recommended.

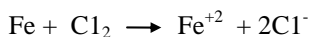
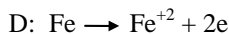
1. Write the redox reaction as 2 half-cell reactions. A half-cell reaction representing oxidation and the other half-cell reaction representing reduction.
2. Balance each half-cell individually.
3. To balance oxygen, add H_2O molecules to the side where oxygen needs to be added.
4. To balance hydrogen add H^+ to the opposite side where water molecules were added.
5. The charge on each half-cell reaction is balanced by adding electrons to the appropriate side (usually on the same side where H^+ are added).
6. To balance the two half-cell reactions we need to make the number of e's lost equal number of e's gained by multiplying the two half-cell reactions by the appropriate number and adding:



Now we balance each half-cell reaction individually.

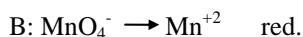
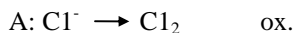
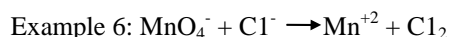


Note that half-cell reaction C gained 2e while half-cell reaction D lost 2e. Therefore, the number of e's lost is equal to the number of e's gained. To get the total reaction we simply add reactions C and D.

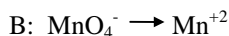
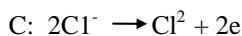


Note: Cl_2 is the oxidation agent; it oxidized Fe from oxidation state of zero to +2. The Fe is the reducing agent; it reduced the oxidation number of chlorine from 0 to -1.

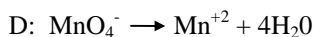
Also, note that the reducing agent (Fe) was oxidized and the oxidizing agent (Cl_2) was reduced.



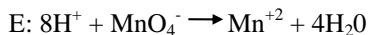
Balance each half-cell reaction. To balance A multiply Cl by 2; then add 2e to the right side of the reaction to obtain reaction C.



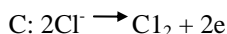
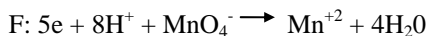
To balance oxygen in half-cell reaction B, add 4 moles of H_2O to the right side to obtain reaction D.



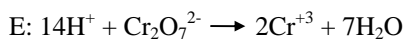
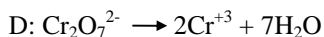
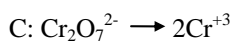
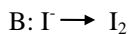
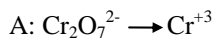
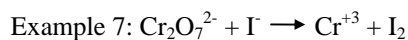
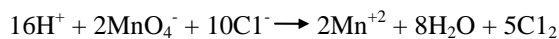
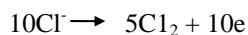
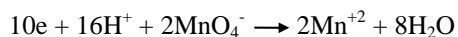
To balance the hydrogen add 8H^+ to the opposite side where H_2O molecules were added and obtain E.



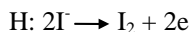
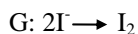
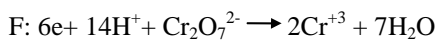
The net charge on the left side of the half-cell reaction E is +7 and on the right side is +2. To balance the charge add 5e to the left side of the reaction to obtain F.



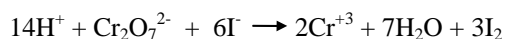
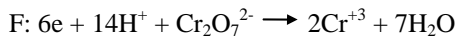
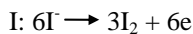
Half-cell reaction F gained 5e, while half-cell reaction C lost 2e. To balance the gain and loss **multiply** half-cell reaction F by 2 and half-cell reaction C by 5 and add.



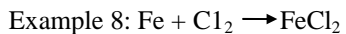
The net charge on the left side of the half-cell reaction is +12 (14 - 2 = 12), while on the right side is +6 (2 x 3). To balance the charge add 6e to the left side and obtain Reaction F.



To balance the gain and the loss of e's, multiply Reaction H by 3 obtaining Reaction I. Add I and F.



BALANCING REDOX REACTIONS BY CHANGE IN OXIDATION NUMBER METHOD:



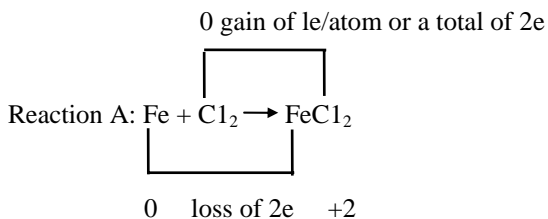
Oxidation number of Fe = 0

Oxidation number of Cl₂ = 0

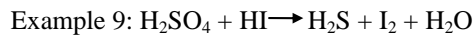
Oxidation number of Fe in FeCl₂ = +2

Oxidation number of Cl in FeCl₂ = -1

The summary is given by Reaction A:



The loss of e's is equal to the gain of e's.



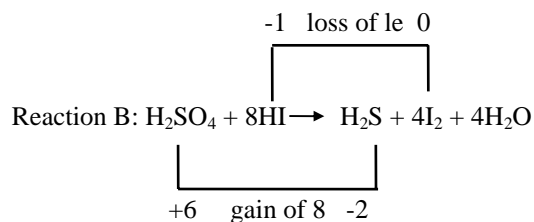
Establish the oxidation number of each atom in the equation.

$$\text{S in H}_2\text{SO}_4 = +6$$

$$\text{I in HI} = -1$$

$$\text{S in H}_2\text{S} = -2$$

The summary is shown by Reaction B.



On the left side we have 8HI; therefore, multiply I₂ on the right side by 4 to balance the I. Left side 4 oxygen; therefore, multiply H₂O on the right side by 4 to balance the oxygen.



Oxidation number of Cl in HCl is -1

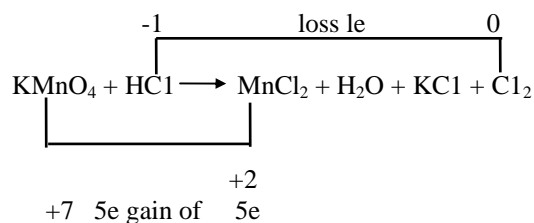
Oxidation number of Mn in KMnO₄ is +7

Oxidation number of Mn in MnCl₂ is +2

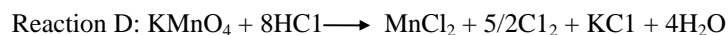
Oxidation number of Cl₂ is 0

The summary is shown by Reaction C

Reaction C:

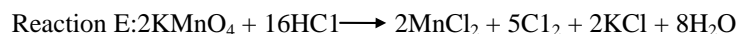


We multiply the HCl by 5, but in this case we have to notice that the HCl was partly used as a reducing agent and partly used to supply the Mn^{+2} and the K^+ with Cl^- . Therefore, in addition to the 5 molecules we need to add 2 molecules of HCl to supply 2Cl^- for the Mn^{+2} and one molecule of HCl to supply one chloride ion for the K^+ . Adding these gives 8 HCl as shown by Reaction D.

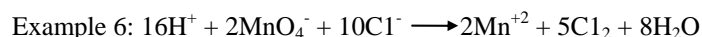


The rest of the reaction is balanced by inspection.

On the left side we have 4 oxygens; therefore, multiply the H_2O by 4. On the left side we have 8 chlorines; therefore, multiply the Cl_2 by $5/2$. We already have 3 other chlorine (MnCl_2 and KCl). To get rid of the fraction, multiply all compounds in reaction D by 2 to obtain Reaction E.



Compare example 6 with example 10; it is seen that both of these examples are the same.



The reason why we have 16H^+ and 10Cl^- in the ion-electron method is because we did not add the 4Cl^- for 2Mn^{+2} nor did we add the Cl^- attached to the K^+ .

The equivalent weight of oxidizing or a reducing agent is equal to its molecular weight divided by the number of transferred e^- 's.

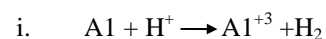
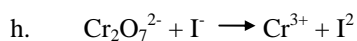
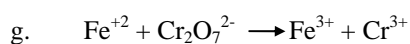
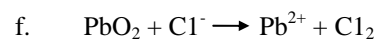
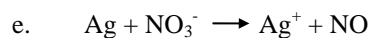
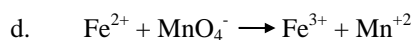
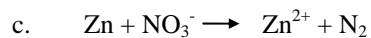
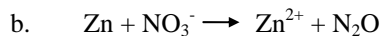
$$\text{Equivalent weight of } \text{KMnO}_4 \text{ in example 10} = \frac{158}{5} = 31.6$$

$$\text{Equivalent weight of } \text{H}_2\text{SO}_4 \text{ in example 9 of } \frac{98}{8} = 12.25$$

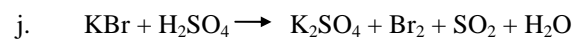
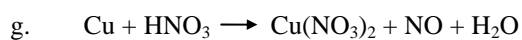
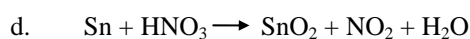
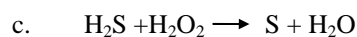
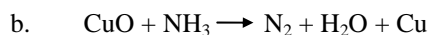
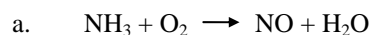
So then to prepare 1N solution of KMnO_4 used in example 10, dissolve 31.6 g of KMnO_4 in 1 liter of solution.

Section II

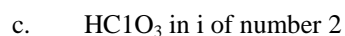
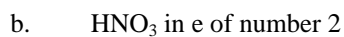
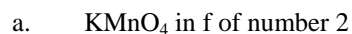
1. Balance the following redox reactions using the ion electron method:



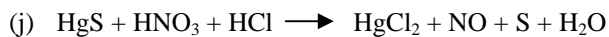
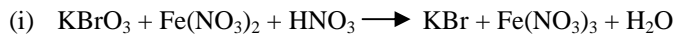
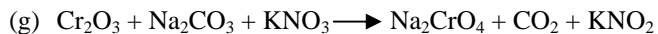
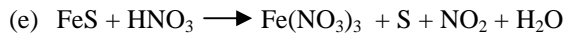
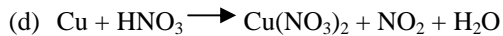
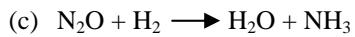
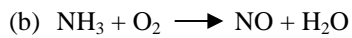
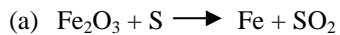
2. Balance the following redox reactions using the change in oxidation number method.



3. Calculate the equivalent weight of:

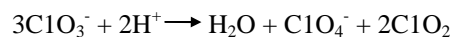
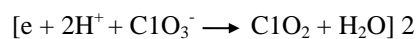
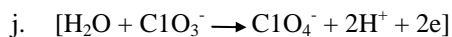


4. Balance the following redox reactions using both the change in oxidation number method and the electron method.

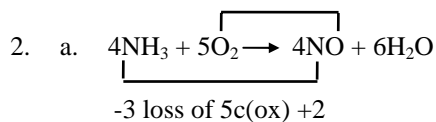


ANSWERS

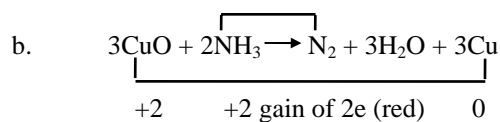
1. a. $\text{Zn} \rightarrow \text{Zn}^{+2} + 2\text{e}$
 $[\text{e} + 2\text{H}^+ + \text{NO}_3^- \rightarrow \text{NO}_2 + \text{H}_2\text{O}] 2$
 $2/ + \text{Zn} + 4\text{H}^+ + 2\text{NO}_3^- \rightarrow \text{Zn}^{+2} + 2\text{NO}_2 + 2\text{H}_2\text{O} + 2/\text{e}$
- b. $[\text{Zn} \rightarrow \text{Zn}^{+2} + 2\text{e}] 4$
 $8\text{e} + 10\text{H}^+ + 2\text{NO}_3^- \rightarrow \text{N}_2\text{O} + 5\text{H}_2\text{O}$
 $4\text{Zn} + 8\text{e} + 10\text{H}^+ + 2\text{NO}_3^- \rightarrow 4\text{Zn}^{+2} + \text{N}_2\text{O} + 5\text{H}_2\text{O} + 8\text{e}$
- c. $[\text{Zn} \rightarrow \text{Zn}^{+2} + 2\text{e}] 5$
 $10\text{e} + 12\text{H}^+ + 2\text{NO}_3^- \rightarrow \text{N}_2 + 6\text{H}_2\text{O}$
 $10\text{e} + 5\text{Zn} + 12\text{H}^+ + 2\text{NO}_3^- \rightarrow 5\text{Zn}^{+2} + \text{N}_2 + 6\text{H}_2\text{O} + 10\text{e}$
- d. $[\text{Fe}^{+2} \rightarrow \text{Fe}^{+3} + \text{e}] 5$
 $5\text{e} + 8\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{+2} + 4\text{H}_2\text{O}$
 $5\text{e} + 5\text{Fe}^{+2} + 8\text{H}^+ + \text{MnO}_4^- \rightarrow 5\text{Fe}^{+3} + \text{Mn}^{+2} + 4\text{H}_2\text{O} + 5\text{e}$
- e. $[\text{Ag} \rightarrow \text{Ag}^+ + \text{e}] 3$
 $3\text{e} + 4\text{H}^+ + \text{NO}_3^- \rightarrow \text{NO} + 2\text{H}_2\text{O}$
 $3\text{Ag} + 3\text{e} + 4\text{H}^+ + \text{NO}_3^- \rightarrow 3\text{Ag}^+ + \text{NO} + 2\text{H}_2\text{O} + 3\text{e}$
- f. $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}$
 $2\text{e} + 4\text{H}^+ + \text{PbO}_2 \rightarrow \text{Pb}^{+2} + 2\text{H}_2\text{O}$
 $2\text{e} + 2\text{Cl}^- + 4\text{H}^+ + \text{PbO}_2 \rightarrow \text{Cl}_2 + \text{Pb}^{+2} + 2\text{H}_2\text{O} + 2\text{e}$
- g. $[\text{Fe}^{+2} \rightarrow \text{Fe}^{+3} + \text{e}] 6$
 $6\text{e} + 14\text{H}^+ + \text{Cr}_2\text{O}_7^{=} \rightarrow 2\text{Cr}^{+3} + 7\text{H}_2\text{O}$
 $6\text{Fe}^{+2} + 6\text{e} + 14\text{H}^+ + \text{Cr}_2\text{O}_7^{=} \rightarrow 6\text{Fe}^{+3} + 2\text{Cr}^{+3} + 7\text{H}_2\text{O} + 6\text{e}$
- h. $[2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}] 3$
 $6\text{e} + 14\text{H}^+ + \text{Cr}_2\text{O}_7^{=} \rightarrow 2\text{Cr}^{+3} + 7\text{H}_2\text{O}$
 $6\text{e} + 14\text{H}^+ + \text{Cr}_2\text{O}_7^{=} + 6\text{I}^- \rightarrow 2\text{Cr}^{+3} + 7\text{H}_2\text{O} + 3\text{I}_2 + 6\text{e}$
- i. $[\text{Al} \rightarrow \text{Al}^{+3} + 3\text{e}] 2$
 $[2\text{e} + 2\text{H}^+ \rightarrow \text{H}_2] 3$
 $2\text{Al} + 6\text{e} + 6\text{H}^+ \rightarrow 2\text{Al}^{+3} + 6\text{e} + 3\text{H}_2$



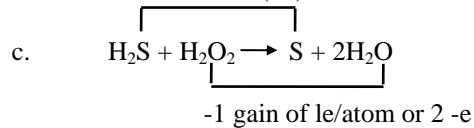
0 gain 2e/at or 4e



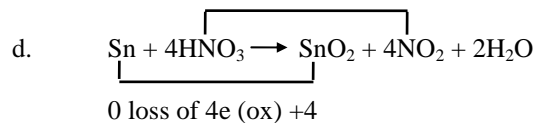
-3 loss of 3e 0(ox)



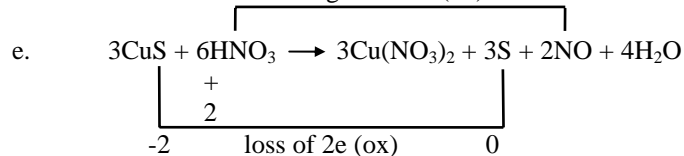
-2 loss of 2e 0(ox)



+5 gain of 1e (red) +4

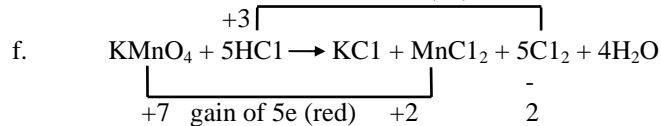


+5 gain of 3e (ox) +2

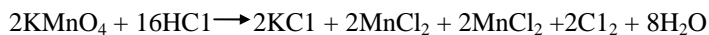


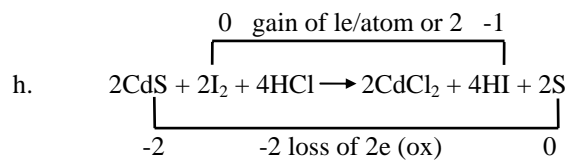
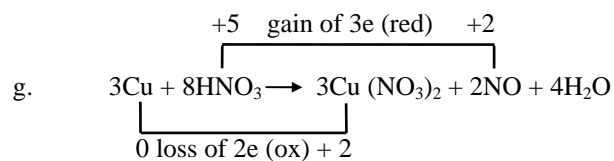
8

" -1 loss of 1e (ox) 0

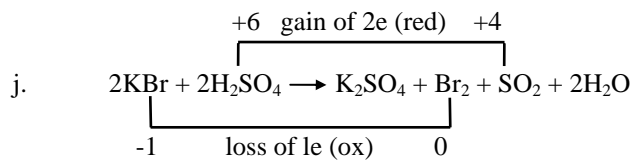
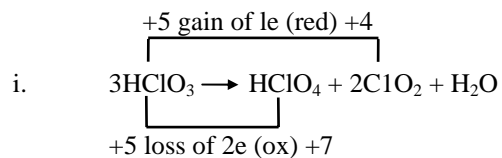
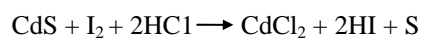


or





or



3. a. Equivalent Weight of $\text{KMnO}_4 = 158.1/5 = 31.62$
 b. Equivalent Weight of $\text{HNO}_3 = 63/3 = 21$
 c. Equivalent Weight of $\text{HClO}_3 = 84.5/3 = 28.17$