Balancing Chemical Equations

(This Chapter is good to revise. Understanding of skills involved is essential)

Chemical Equations

The chemical change in a reaction is represented by an equation of formulae known as chemical equation. A balanced chemical equation gives information about
(i) the number of moles of reactants and products involved in the reaction
(ii) the mass of substances reacting and the mass of products formed
(iii) the volume of the gases involved in the reaction at STP (from Avogradro's law, 1 mole of any gas at STP has a volume equal to 22.4 litres i.e. molar volume).
In a chemical reaction, (as different from a nuclear reaction) atoms (and not the molecules or ions) and mass are conserved i.e. sum of the masses of all reactants are equal to the sum of the masses of all products. However, this mass relation does not tell us about variable conditions i.e. whether H_2O in a reaction is liquid or steam.
A chemical equation does not give us information about
(i) physical status of reactants and products
(ii) heat changes in the reaction
(iii) conditions, rate and mechanism of the reaction
A balanced chemical equation correlates the masses of the products with those of the reactants. This balance sheet of masses determined by chemical equations and formulae is called Stoichiometry and may be used to correlate and calculate amounts of reactants and products in a reaction.
A balanced chemical equation must be balanced with respect to mass as well as charge.

Balancing of Equations

For the purpose of balancing, the equations may be grouped into two categories:
(a) Simple equations
(b) Redox equations.

Rules for balancing simple equations

Balancing by inspection
(i) Balance the metals first. If many metals are present, balance them one by one.
(ii) Balance the non-metals except hydrogen and oxygen. If many non-metals are present, balance them one by one.
(iii) Balance hydrogen and oxygen.
(iv) Check all the elements and if any thing is still not balanced, repeat the same order till all the elements are balanced.

Example

Balance the equation:
$Na_2Cr_2O_7 + HCl \rightarrow CrO_2Cl_2 + NaCl + H_2O$
Solution
(i) Balance Na and Cr
$Na_2Cr_2O_7 + HCl \rightarrow 2CrO_2Cl_2 + 2NaCl + H_2O$
(ii) Balance Cl
$Na_2Cr_2O_7 + 6\ HCl \rightarrow 2CrO_2Cl_2 + 2NaCl + H_2O$
(iii) Balance H and O
$Na_2Cr_2O_7 + 6\ HCl \rightarrow 2CrO_2Cl_2 + 2NaCl + 3H_2O$
(iv) Check whether all elements are balanced or not.

Example

Balance the equation:

$$Zn_3As_2 + H_2SO_4 \rightarrow ZnSO_4 + AsH_3$$

Solution

(i) Balance Zn and As

$$Zn_3As_2 + H_2SO_4 \rightarrow 3ZnSO_4 + 2AsH_3$$

(ii) Balance S

$$Zn_3As_2 + 3H_2SO_4 \rightarrow 3ZnSO_4 + 2AsH_3$$

(iii) Balance H and O

(iv) Check whether all elements are balanced or not.

Balancing Redox equations

Equations in which reactants are both oxidised and reduced are called redox reactions. In these reactions, both oxidation and reduction takes place.

In reactions of ionic compounds, it is easier to ascertain loss or gain of electrons. In other types of compounds, to determine whether oxidation or reduction has taken place, it is convenient to use the concept of oxidation number.

Oxidation Number

Oxidation number of an element in a compound indicates the amount of oxidation or reduction which is required to convert its one atom from free state to that state in which it is present in the compound. If oxidation is required then oxidation number is positive and if reduction is required then oxidation number is negative.

In applying the concept of oxidation number to covalent compounds, a shared pair of electrons between two different atoms is believed to be present at more electronegative element out of the two. Remember that in a shared pair there is no loss or gain of electrons but electrons are assigned to more electronegative atom. Thus oxidation number is the electrical charge assigned to a particular atom according to some arbitrary rules based on electronegativity. Since it is a concept, oxidation number can also be fractional though fractional transfer of electrons is not possible.

Rules for determining oxidation number

1. Oxidation number of free or uncombined atom is zero.
2. Oxidation number of an ion is equal to the charge on it. Thus oxidation number of Al in Al^{3+} is +3 and that of S in S^{2-} is -2.
3. The algebraic sum of oxidation number of all the atoms in a neutral molecule is equal to zero.
4. The algebraic sum of oxidation number of all the atoms in an ion is equal to the charge on that ion. Thus, oxidation number of SO_4^{2-} is -2.
5. The oxidation number of alkali metals of group IA (new group number 1) in their compounds is +1.
6. The oxidation number of alkaline earth metals of group IIA (new group number 2) in their compounds is +2.
7. The oxidation number of fluorine in all its compounds is -1.
8. The oxidation number of hydrogen in its compound is generally +1 except in metal hydrides of group IA and group IIA like NaH, CaH_2 where the oxidation number of H is -1.
9. The oxidation number of oxygen in its compound is generally -2, except in peroxides like H_2O_2 where it is -1, superoxides like KO_2 where it is -1/2 and in F_2O

where it is +2.
10. The oxidation number of halogens (Cl, Br and I) in their compounds is generally -1, except when they combine with F, O or a more electronegative atom.
11. Elements may have different oxidation numbers in different compounds.
12. Oxidation number of an element in a compound is generally not more than its group number.

Note:
Though often used interchangeably, the term 'oxidation state' is actually different from 'oxidation number'. Oxidation number, as we have seen, is a result of 'assigning' electrons (sometimes arbitrarily) to an atom and is used only for the purpose of balancing redox equations. Oxidation state is not (at least normally) arbitrary. The oxidation number may be fractional but oxidation state is always a whole number.

Learn to determine oxidation number

Example
Determine the oxidation number of
(i) Mn in $KMnO_4$ (ii) Cr in $K_2Cr_2O_7$
(iii) Cl in ClO^-_4 (iv) N in N_3H
(v) S in $H_2S_2O_8$ (vi) Fe in $Fe_{0.9}O$
Solution
(i) Oxidation number of K = +1, O = -2
Hence, + 1 + Mn + 4 (-2) = 0
$\therefore$ Mn = +7
(ii) Oxidation number of K = +1, O = -2
Hence, 2(+1) + 2 Cr + 7 (-2) = 0
$\therefore$ Cr = + 6
(iii) Oxidation number of O = -2
Hence, Cl + 4 (-2) = -1
$\therefore$ Cl = +7
(iv) Oxidation number of H = +1
Hence, +1 + 3N = O

$$\therefore N = -\frac{1}{3}$$

(In fact, in HN_3, the oxidation number of the different N atoms are different)

$$\begin{array}{lcccc} & H - & \overset{-}{N} - & \overset{+}{N} \equiv & N \\ \text{Oxidation Number} & +1 & -1 & 0 & 0 \end{array}$$

When we take the average of the three nitrogen atoms, it becomes $-\frac{1}{3}$.
(v) Oxidation number of H = +1, O = -2
Hence 2(+1) + 2S + 8(-2) = 0
or, S = + 7, which is not possible as S is in VI group.
So, the oxidation number of S can maximum be +6. The anomaly is because of the fact that it has peroxide linkage where the two oxygen have an oxidation number equal to -1, whereas, we made the calculation on the basis of both having oxidation number -2.
Now, 2(+1) +2S + 2(-1) + 6(-2) = 0
and S = + 6

(vi) $Fe_{0.9}O$ is a non-stoichiometric compound because atoms do not combine in fractions, and therefore, simple rules will not be applicable in this example. Obviously, the compound has metal deficiency because of the presence of some Fe^{3+} ions together with Fe^{2+} ions.
Hence, oxidation states of Fe in such a compound is both +2 and +3 (and not 2/0.9)

Balancing of redox equations
(1) Oxidation number method
Step I
Write oxidation number of each element in the equation.
Step II
Indicate the changes in oxidation number and find out oxidation and reduction and their magnitudes. (Note: Increase in oxidation number is oxidation and decrease in oxidation number is reduction.)
Step III
Multiply the substances oxidised and reduced by proper coefficients to make the increase in oxidation number equal to the decrease i.e. make oxidation = reduction.
Step IV
Balance the simple equation, balancing metals first, then the non-metal, then H and O and check. If not balanced, repeat Step IV.

1. Balancing a simple redox equation

Example
Balance the equation:
$K_2CrO_4 + Cu_2O \rightarrow Cu(OH)_2 + K[Cr(OH)_4] + KOH + H_2O$

Solution

Decrease, 3 per molecule

$$K_2\overset{6+}{Cr}O_4 + \overset{1+}{Cu_2}O \rightarrow \overset{2+}{Cu}(OH)_2 + K[Cr^{3+}(OH)_4] + KOH$$

Increase, 1 per atom or 2 per molecule

Decrease per oxidant $[(K_2CrO_4) = 3] \times 2 = 6$
Increase per reductant $[(Cu_2O) = 2] \times 3 = 6$
Write LHS and make adjustment for metals (K, Cu, Cr) on RHS.

$$\boxed{2}\,K_2CrO_4 + \boxed{3}\,Cu_2O + H_2O \rightarrow \boxed{6}\,Cu(OH)_2 + \boxed{2}\,K[Cr(OH)_4] + \boxed{2}\,KOH$$

We may now balance either with respect to H or with respect to O using H_2O. Since O occurs in more than one molecule in LHS, using H is easier.
In RHS, we have 12 + 8 + 2 = 22 H
We add 11 H_2O in LHS $2K_2CrO_4 + 3Cu_2O + \boxed{11}\,H_2O \rightarrow 6Cu(OH)_2 + 2K[Cr(OH)_4] + 2KOH$
Check again that equation is balanced.

Note:
We made the final adjustment using water(H_2O). Usually after adjusting the ratio of

reductant and oxidant, we have to adjust using some acid or base and/or H_2O to do the final balancing. The situation becomes complicated when the reductant or oxidant is also an acid or base i.e. it plays a dual role.

2. When reductant plays a dual role

Example

Balance the equation:

$CuF_2 + NH_3 \rightarrow Cu_3N + NH_4F + N_2$

Step I $\quad \overset{2+}{Cu}\,\overset{1-}{F_2} + \overset{3-}{N}\,\overset{1+}{H_3} \rightarrow \overset{1+}{Cu_3}\,\overset{3-}{N} + \overset{3-}{N}\,\overset{1+}{H_4}\,\overset{1-}{F} + \overset{0}{N_2}$

Step II

Oxidation, Increase of 3 per molecule

$$\overset{2+}{Cu}\,F_2 + \overset{3-}{N}\,H_3 \rightarrow \overset{1+}{Cu_3}\,N + NH_4F + \overset{0}{N_2}$$

Reduction, decrease of 1 per molecule

Note:

The oxidation numbers which do not change are not written on the product side, as they are written on the reactant side. Oxidation numbers of H(1+) and F(1-) do not change. Similarly N(3-) is same in NH_3, Cu_3N and NH_4F. So we have not written these.

Step III

In order to make the increase in oxidation number = decrease in oxidation number, we must have three molecules of CuF_2 for every molecule of NH_3. We should thus multiply CuF_2 by 3 and NH_3 by 1.

The product has N_2 and Cu_3N. $3CuF_2$ can give Cu_3N but one NH_3 will give only $1/2N_2$. To get a whole number coefficient, we should have an even number as coefficient of NH_3. So, we maintain a 3:1 ratio for CuF_2 and NH_3. Multiply CuF_2 by 6 and NH_3 by 2.

$$\boxed{6}\,CuF_2 + \boxed{2}\,NH_3 \rightarrow Cu_3N + NH_4F + N_2$$

unbalanced

Step IV

NH_3 has two roles to play - one is to reduce $Cu^{2+} \rightarrow Cu^{1+}$ and the other is to produce Cu_3N and NH_4F. The coefficient of NH_3 shown takes care of only the redox part. We, therefore, have to add extra NH_3. We will, however, do it later.

Balance metals first

$$6CuF_2 \rightarrow 2Cu_3N$$

Balance F

$$6CuF_2 \rightarrow 12NH_4F$$

$\therefore 6CuF_2 + 2\,NH_3 \rightarrow 2\,Cu_3N + 12\,NH_4F + N_2$

Balance N.

On right hand side, we have $2(Cu_3N) + 12(NH_4F) = 14$ N atoms where the oxidation number is the same as in NH_3(3-). So on the left hand side we add $14NH_3$. Thus we will have $2NH_3$ (for reduction) + $14NH_3$ (no redox) = $16NH_3$.

Rewrite equation:

$6CuF_2 + \boxed{16}\, NH_3 \rightarrow 2Cu_3N + 12NH_4F + N_2$

The reaction is now balanced.
H is automatically balanced.

3. When the oxidant plays a dual role

Example
Balance the equation:

$Cu + HNO_3 \rightarrow Cu(NO_3)_2 + H_2O + NO$

Solution

decrease, 3 per molecule

$\overset{0}{Cu} + H\overset{5+}{N}O_3 \rightarrow \overset{2+}{Cu}(NO_3)_2 + H_2O + \overset{2+}{N}O$

Increase, 2 per atom

Decrease per oxidant (HNO_3) = 3 } $\times$ 2 = 6
Increase per reductant (Cu) = 2 } $\times$ 3 = 6
Write LHS and make adjustments on RHS.

$\boxed{3}Cu + \boxed{2}\, HNO_3 \rightarrow \boxed{3}Cu(NO_3)_2 + H_2O + \boxed{2}\, NO$

HNO_3 plays a dual role. It oxidises Cu as an oxidant and it also provides NO_3^- ions as an acid to produce $Cu(NO_3)_2$. So we have to make final adjustments by adding required number of HNO_3 to what is required as an oxidant.

In RHS, we have, $3Cu(NO_3)_2 \longrightarrow 6NO_3^-$
$\Rightarrow$ We must add $6HNO_3$ in LHS to have 2 + 6 = $8HNO_3$. Rewrite and adjust for H using H_2O in the R.H.S.

$3\,Cu + \boxed{8}\, HNO_3 \rightarrow 3\,Cu(NO_3)_2 + \boxed{4}\, H_2O + 2NO$

The equation is balanced.

In an earlier example, we have already seen that we have to consider the change in oxidation number per molecule of oxidant or reductant (and not per atom) to balance the equation. Sometimes an oxidant or reductant may have two different elements undergoing reduction or oxidation. The strategy, however, remains the same. We consider change in oxidation number per molecule.

4. When oxidant or reductant has more than one element by changing oxidation number

Example
Balance the equation:

$FeO.Cr_2O_3 + Na_2CO_3 + O_2 \longrightarrow Na_2CrO_4 + Fe_2O_3 + CO_2$

Solution
Write the oxidation numbers which change and indicate total changes in oxidation

number per molecule.

Decrease, 2 per atom or 4 per molecule

$$\overset{2+}{Fe}O.\overset{3+}{Cr_2}O_3 + NaCO_3 + \overset{0}{O_2} \longrightarrow Na_2\overset{6+}{Cr}\overset{2-}{O_4} + \overset{3+}{Fe_2}\overset{2-}{O_3} + C\overset{2-}{O_2}$$

Increase, 3 per atom or 6 per molecule

Increase, 1 per atom

Total increase per $FeO.Cr_2O_3$ = 6 + 1 = 7 $\times$ 4 = 28
Total decrease per O_2 = 4 x 7 = 28
Write LHS and make necessary changes on RHS

$$\boxed{4}FeO.Cr_2O_3 + NaCO_3 + \boxed{7}O_2 \rightarrow \boxed{8}Na_2\ CrO_4 + \boxed{2}\ FeO_3 + CO_2$$

Now also balance for Na. RHS has 8 $\times$ 2 = 16Na
So in LHS we require $8Na_2CO_3$. Make necessary changes on RHS.
Rewrite

$$4\,FeO.Cr_2O_3 + \boxed{8}Na_2CO_3 + 7O_2 \rightarrow 8Na_2CrO_4 + 2Fe_2O_3 + \boxed{8}CO_2$$

Equation is balanced.

Example
Balance the equation:
$CrI_3 + KOH + Cl_2 \rightarrow K_2CrO_4 + KIO_4 + KCl + H_2O$
Solution

decrease, 1 per atom or 2 per molecule

$$\overset{3+}{Cr}\overset{1-}{I_3} + KOH + \overset{0}{Cl_2} \longrightarrow K_2\overset{6+}{Cr}O_4 + K\overset{7+}{I}O_4 + K\overset{1-}{Cl} + H_2O$$

increase, 8 per atom or 24 per molecule

increase, 3 per molecule

Total increase per reductant (CrI_3) = 24+3 = 27 $\times$ 2 = 54
Total decrease per oxidant (Cl_2) = 2 $\times$ 27 = 54
Write LHS and make necessary changes on RHS.

$$\boxed{2}CrI_3 + KOH + \boxed{27}Cl_2 \rightarrow \boxed{2}K_2CrO_4 + \boxed{6}KIO_4 + \boxed{54}KCl + H_2O$$

Balance K (4 + 6 + 54 = 64) of RHS by using 64KOH on LHS and adjust H with H_2O on RHS.

$$2CrI_3 + \boxed{64}KOH + 27Cl_2 \rightarrow 2K_2CrO_4 + 6KIO_4 + 54KCl + \boxed{32}H_2O$$

Equation is balanced.

5. When assigning of oxidation number is ambiguous

The situation may become complicated when the oxidation states of more than one element undergoes change. Some of the complications are discussed in the following example.

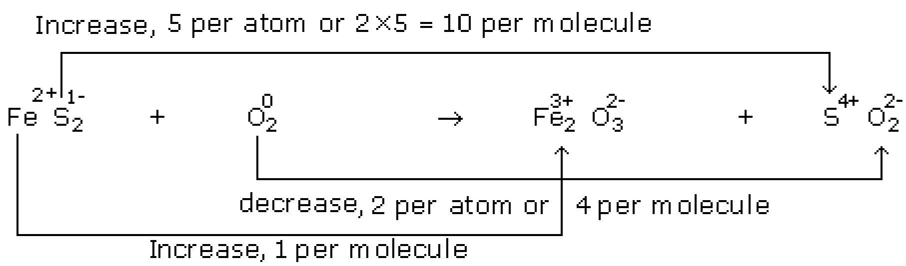
Example
Balance $FeS_2 + O_2 \rightarrow Fe_2O_3 + SO_2$

Solution

This apparently simple looking reaction may provide several conceptual problems such as

- What are the oxidation states of Fe and S in FeS_2?
  There are two possibilities (i) Fe^{2+} and S^{1-}
  (ii) Fe^{4+} and S^{2-}

  We may use any one of the possibilities - both are self consistent.
  Case (i) means during the reaction, the oxidation number of both Fe and S will increase.
  Case (ii) means that the oxidation number of Fe will decrease and that of S will increase.
  How do we handle case (ii)?
  In the case (ii) we will consider the net increase or decrease of oxidation number per molecule of FeS_2.
- Both the possibilities for assignment of oxidation number are unrealistic oxidation states. What are the oxidation states?
  Actually iron pyrite contains a combination of one S^{2-} and one S (S^{2-} + S = S_2^{2-})
  Balancing using case (i)

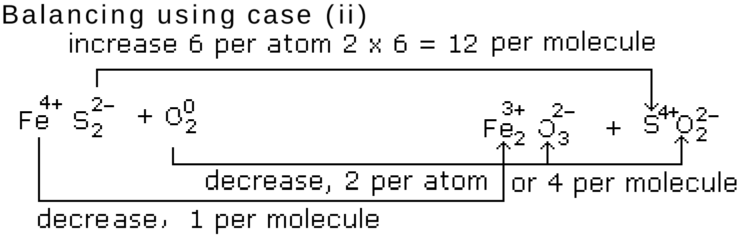


Total increase per molecule (FeS_2) = 10 + 1 = 11} × 4 = 44
Total decrease per molecule (O_2) = 4} × 11 = 44
Multiply FeS_2 by 4 and O_2 by 11 and adjust RHS.

$\boxed{4}FeS_2 + \boxed{11}O_2 \rightarrow \boxed{2}Fe_2O_3 + \boxed{8}SO_2$

Equation is balanced.

Balancing using case (ii)



Total increase per molecule (FeS_2) = 12 -1 = 11} × 4 = 44
Total decrease per molecule (O_2) = 4} × 11 = 44
The rest is same as above.
Note: Both ways we end up with $4FeS_2 + 11O_2$.

The case (ii) above is an example where we consider some atoms increasing oxidation number and some atoms decreasing oxidation number in the same molecule. A complicated situation arises when some atoms get oxidised and some get reduced, for the same element. Such reactions are called disproportionation reactions.

Balancing disproportionation reactions

Example

$Br_2 + NaOH \rightarrow NaBr + NaBrO_3 + H_2O$

Solution

Here, some atoms of Br in Br_2 are oxidised to $NaBrO_3$ and some atoms of Br are reduced to NaBr. How do we adjust the ratio of oxidant to reductant? It may be done in the usual way but a simple trick makes it easier - as shown below.
The trick we use is to calculate the oxidation number that changes in the reverse way i.e. RHS to LHS.

increase 1 per molecule

$$\overset{0}{Br_2} + NaOH \longrightarrow Na\overset{-1}{Br} + Na\overset{5+}{Br}O_3 + H_2O$$

decrease, 5 per molecule

Thus from RHS $\rightarrow$ LHS

Total increase per molecule (NaBr) = 1} $\times$ 5 = 5
Total decrease per molecule ($NaBrO_3$) = 5} $\times$ 1 = 5
We multiply NaBr by 5 and $NaBrO_3$ by 1 in RHS and adjust LHS (for Br and Na)

$$\boxed{3}\ Br_2 + \boxed{6}NaOH \rightarrow \boxed{5}NaBr + NaBrO_3 + H_2O$$

Now adjust H by adding H_2O in RHS.

$$3\ Br_2 + 6\ NaOH \rightarrow 5\ NaBr + NaBrO_3 + \boxed{3}H_2O$$

Example

Balance $P_4 + NaOH + H_2O \rightarrow PH_3 + NaH_2PO_3$

Solution

increase, 3 per molecule

$$\overset{0}{P_4} + NaOH \longrightarrow \overset{3-}{P}H_3 + NaH_2\overset{1+}{P}O_2 + H_2O$$

decrease, 1 per molecule

Thus on RHS we should have PH_3:Na H_2PO_2 in the ratio 1:3. Using the method shown above we have

$$P_4 + \boxed{3}NaOH + H_2O \rightarrow PH_3 + \boxed{3}NaH_2PO_2$$

Adjusting H, using H_2O in LHS

$$P_4 + 3NaOH + \boxed{3}H_2O \rightarrow PH_3 + 3NaH_2PO_2$$

Equation is balanced

Ion-electron or Half equation method for balancing equations

Most reactions involving aqueous solutions also involve ions. Equations of such reactions may be balanced using the ion electron method . The final balanced equation may come in the form of a net ionic equation or the full equation may be reconstituted. The net ionic equation must be balanced with respect to charge as well as with respect to mass.

The method

Step I
Write the unbalanced net ionic equation.

Step II
Divide the reaction into two half reactions - one for oxidation and the other for reduction. In case of doubt use concept of oxidation number to find which species is being oxidised and which species is being reduced. Note that calculating oxidation number is not necessary for this method.

Step III
Balance mass in each half equation separately.
In acidic medium add H^+ and H_2O wherever necessary to balance H and O. Normally add double the H^+ to the side rich in oxygen and add H_2O molecules equal to excess oxygen to the opposite side.
In basic and neutral medium balance H and O by adding OH^- and H_2O. In general, add equal number of H_2O to side rich in oxygen and double OH^- to the opposite side.

Step IV
Balance the charge by adding electrons to more positive side. In general, in oxidation half-reaction, electrons will appear on the right hand side while in reduction half reaction electrons will appear on the left hand side.

Note:
The number of electrons will be equal to the change in oxidation number in both cases. This may be used to check the correctness of balanced equation.)

Step V
Multiply half equations with suitable coefficients to make number of electrons equal in both half equations as electrons do not appear in the final equation.

Step VI
Restoration to molecular equation (optional): Add both half equations and then add suitable cations and anions in equal numbers to both sides wherever required.

Example
Balance
$FeSO_4 + KMnO_4 + H_2SO_4 \rightarrow Fe_2(SO_4)_3 + MnSO_4 + K_2SO_4 + H_2O$

Solution
Step I
$Fe^{2+} + SO_4^{2-} + K^+ + MnO_4^- + H^+ \rightarrow Fe^{3+} + SO_4^{2-} + Mn^{2+} + K^+$
Step II

Oxidation	Reduction
$Fe^{2+} \rightarrow Fe^{3+}$	$MnO_4^- \rightarrow Mn^{2+}$
Oxidation number: + 2 + 3	+ 7 + 2

Step III
(Acidic Medium) Mass balance

$Fe^{2+} \rightarrow Fe^{3+}$

$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$

Step IV
Charge balance

$Fe^{2+} \rightarrow Fe^{3+} + 1e^-$

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

Step V
Multiply by suitable coefficient

$5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

Step VI

$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$

Final equation

$$5FeSO_4 + KMnO_4 + 4H_2SO_4 \rightarrow \frac{5}{2}Fe_2(SO_4)_3 + MnSO_4 + 4H_2O + \frac{1}{2}K_2SO_4$$

Multiply by 2 to get final balanced equation

$$10FeSO_4 + 2KMnO_4 + 8H_2SO_4 \rightarrow 5Fe_2(SO_4)_3 + 2MnSO_4 + 8H_2O + K_2SO_4$$

Example
Balance

$ClO_2 + KSbO_2 \rightarrow KClO_2 + KSb(OH)_6$ in basic medium

Solution
Step I

$ClO_2 + K^+ + SbO_2^- \rightarrow K^+ + ClO_2^- + K^+Sb(OH)_6^-$

Step II

Oxidation	Reduction
$SbO_2^- \rightarrow Sb(OH)_6^-$	$ClO_2 \rightarrow ClO_2^-$

Step III
Mass balance in basic medium.

$SbO_2^- + 2H_2O + 2OH^- \rightarrow Sb(OH)_6^-$

$ClO_2 + 1e^- \rightarrow ClO_2^-$

Step IV
Charge balance

$$SbO_2^- + 2H_2O + 2OH^- \rightarrow Sb(OH)_6^- + 2e^-$$

$$ClO_2 + 1e^- \rightarrow ClO_2^-$$

Step V

Multiply by suitable coefficients

$$SbO_2^- + 2H_2O + 2OH^- \rightarrow Sb(OH)_6^- + 2e^-$$

$$2ClO_2 + 2e^- \rightarrow 2ClO_2^-$$

Step VI

$$\underset{K^+}{SbO_2^-} + 2H_2O + \underset{2K^+}{2OH^-} + 2ClO_2 \rightarrow \underset{K^+}{Sb(OH)_6^-} + \underset{2K^+}{2ClO_2^-}$$

Final balanced equation

$$KSbO_2 + 2H_2O + 2KOH + 2ClO_2 \rightarrow KSb(OH)_6 + 2KClO_2$$