SCH 4U

BALANCING REDOX REACTIONS using HALF REACTIONS (in acidic conditions)

- REDOX reactions are more complex to balance than standard reactions. Balancing by inspection will not yield a balanced reaction.

- A systematic method for balancing REDOX reactions making use of ***half-reactions*** is known as balancing by the ION-ELECTRON METHOD.

When balancing a reaction, BOTH ***ATOMS*** and ***CHARGES*** must be balanced.

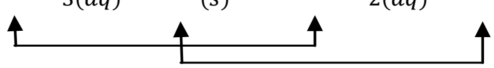
- To balance ***atoms***, ***use coefficients*** (in front of substances).
- To balance ***charges***, ***add electrons*** to the appropriate side of a half-reaction.

EXAMPLE ①: Balance the reaction $CoCl_{3(aq)} + Au_{(s)} \rightarrow CoCl_{2(aq)} + AuCl_{3(aq)}$

STEP 1: Check for the **REDOX** reaction first.

O.N.=

$CoCl_{3(aq)} + Au_{(s)} \rightarrow CoCl_{2(aq)} + AuCl_{3(aq)}$



STEP 2: Write the **NET-IONIC EQUATION** – showing only those ions/molecules that are involved in the reaction. Ignore ***spectator ions*** with unchanged oxidation state.

$Co^{3+}{}_{(aq)} + 3Cl^{-}{}_{(aq)} + Au_{(s)} \rightarrow Co^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)} + Au^{3+}{}_{(aq)} + 3Cl^{-}{}_{(aq)}$

$Co^{3+}{}_{(aq)} + Au_{(s)} \rightarrow Co^{2+}{}_{(aq)} + Au^{3+}{}_{(aq)}$

O.N. of Co has decreased from 3+ to 2+, so it is reduced and is called the oxidizing agent.
O.N. of Au has increased from 0 to 3+, so it is oxidized and is called the reducing agent.

STEP 3: Separate the reaction into **half-reactions**, pairing reactant ion with product ion.

The half reactions are:

$Co^{3+} \rightarrow Co^{2+}$ **Number of particles** are balanced in each half reaction.

$Au \rightarrow Au^{3+}$ However, **charges** are not balanced, so electrons are added to the appropriate side of each equation.

STEP 4: **Add electrons** to appropriate side of each half-reaction to balance charges.

$$e^- + Co^{3+} \rightarrow Co^{2+}$$

$$Au \rightarrow Au^{3+} + 3e^-$$

The number of electrons lost in a half-reaction must be gained by the other half-reaction.
In this case, for each gold atom that loses 3 electrons, 3 of the other half-reaction must occur.
We multiply the first half-reaction by 3.

STEP 5: Multiply 1 or both half-reactions by values that **equalize the number of e$^-$.**
Then ADD (and reduce) the half-reactions together.

$$3 \times (e^- + Co^{3+} \rightarrow Co^{2+})$$

$$3e^- + 3Co^{3+} \rightarrow 3Co^{2+}$$

$$Au \rightarrow Au^{3+} + 3e^-$$

$$3e^- + 3Co^{3+} + Au \rightarrow 3Co^{2+} + Au^{3+} + 3e^-$$

Add the 2 half-reactions. Reduce parts that are common to both sides.

$$3Co^{3+} + Au \rightarrow 3Co^{2+} + Au^{3+}$$

STEP 6: Refresh the **original equation** using coefficients of combined half reactions.

$$3CoCl_3 + Au \rightarrow 3CoCl_2 + AuCl_3$$

EXAMPLE ②: Balance the reaction $Zn_{(s)} + Fe_2(SO_4)_{3(aq)} \rightarrow ZnSO_{4(aq)} + Fe_{(s)}$.

REDOX REACTIONS in ACIDIC CONDITIONS

- REDOX reactions are carried out in aqueous solutions – H_2O molecules and its ions play important roles in the reaction, depending on the nature of the solution.

- In a REDOX reaction, the acidity or basicity of the solution changes due to the consumption or production of H^+ ions or OH^- ions or water molecules.

- REDOX reactions are generally carried out in solutions containing either excess acid or base, and this information must be known before applying the ion-electron method.

BALANCING REDOX EQUATIONS FOR ACIDIC SOLUTIONS

STEP 1: **Divide the equation into two half-reactions.**

STEP 2: **Balance atoms other than H and O.**

STEP 3: **Balance O by adding H_2O.**

STEP 4: **Balance H by adding H^+.**

STEP 5: **Balance net charge by adding e^-.**

STEP 6: **Make e^- gain equal e^- loss; then add half-reactions.**

STEP 7: **Reduce substances common to both sides of reaction.**

EXAMPLE ③: Balance the equation $MnO_4^{-} + H_2SO_3 \rightarrow SO_4^{2-} + Mn^{2+}$. The reaction occurs in an **acidic** solution.

STEP 1: **$MnO_4^{-} \rightarrow Mn^{2+}$**
$H_2SO_3 \rightarrow SO_4^{2-}$

STEP 2: All atoms except for H and O are already balanced.

STEP 3: $MnO_4^{-} \rightarrow Mn^{2+}$ **$\underline{+\ 4\ H_2O}$**
$\underline{H_2O\ +}$ $H_2SO_3 \rightarrow SO_4^{2-}$

STEP 4: **$\underline{8\ H^{+}\ +}$** $MnO_4^{-} \rightarrow Mn^{2+} + 4\ H_2O$
$H_2O + H_2SO_3 \rightarrow SO_4^{2-}$ **$\underline{+\ 4\ H^{+}}$**

STEP 5: **$\underline{5\ e^{-}\ +}$** $8\ H^{+} + MnO_4^{-} \rightarrow Mn^{2+} + 4\ H_2O$
$H_2O + H_2SO_3 \rightarrow SO_4^{2-} + 4\ H^{+}$ **$\underline{+\ 2\ e^{-}}$**

STEP 6: **$\underline{2\ x}$** $[\ 5\ e^{-} + 8\ H^{+} + MnO_4^{-} \rightarrow Mn^{2+} + 4\ H_2O\ ]$
$\underline{5\ x}$ $[\ H_2O + H_2SO_3 \rightarrow SO_4^{2-} + 4\ H^{+} + 2\ e^{-}\ ]$

$10e^{-} + 16H^{+} + 2MnO_4^{-} + 5H_2O + 5H_2SO_3 \rightarrow 2Mn^{2+} + 8H_2O + 5SO_4^{2-} + 20H^{+} + 10e^{-}$

STEP 7: $2MnO_4^{-} + 5H_2SO_3 \rightarrow 2Mn^{2+} + 3H_2O + 5SO_4^{2-} + 4H^{+}$

NOTICE that each side of the balanced equation has the same number of each type of atom and the same net charge.

Balance each REDOX reaction in acidic conditions.

EXAMPLE ④: $S_{(s)} + HNO_{3(aq)} \rightarrow SO_{2(g)} + NO_{(g)} + H_2O_{(l)}$

EXAMPLE ⑤: $HClO_{2(aq)} + I^{-}{}_{(aq)} \rightarrow Cl_{2(g)} + HIO_{(aq)}$