Subject Code Chem 1 **General Inorganic Chemistry**
Module Code 2.0 **Nomenclature of Inorganic Compounds**
Lesson Code 2.7 **Oxidation and Reduction Reactions Part II**
Time Frame **30 minutes**

| **Components** | **Tasks** | **TA[1] (min)** | **ATA[2] (min)** |
|---|---|---|---|
| **Target**<br> | After working on this module, you are expected to:<br>1. Balance redox equations using oxidation number method and half reaction method. | 1 | |
| **Hook**<br> | In our previous topics, we dealt with the topics on balancing chemical equations. Balancing chemical reaction is important so that we will not deviate from the atomic laws especially the "Law of Conservation of Mass". This means that atoms in a chemical reaction is neither created nor destroyed that is why we need to account every atoms that are present in a chemical reaction. Balancing chemical reactions includes all types of chemical reactions including redox reactions.<br><br>In the previous modules, we learned what a redox reaction is. It is very important to realize that the process of giving and losing electrons is simultaneous in the process of redox reactions. This chemical reaction is always paired, which means that reduction cannot occur without oxidation. This also means that balancing redox reactions is also a means of balancing these changes in electrons. | 2 | |
| **Ignite**<br> | Two methods used in balancing redox reactions are **oxidation number method** and **half-reaction method (ion-electron method).**<br><br>First we will discuss how to balance using the oxidation number method<br><br>**Oxidation number method** is a way of balancing the electrons in a Redox reaction. This method is especially useful if we are not dealing with redox reactions in aqueous solutions or dissolved in either basic or acidic medium<br><br>**Steps in balancing using OXIDATION NUMBER METHOD: (Silberberg, 2006)**<br><br>***Step 1.*** Assign **oxidation numbers** to all elements in the reaction.<br>***Step 2.*** From the changes in oxidation number, identify the | 20 | |

[1] Time allocation suggested by the teacher.
[2] Actual time allocation spent by the student (for information purposes only).

**oxidized and reduced species**.

***Step 3.*** Compute the number of **electrons lost in the oxidation** and **gained in reduction** from the oxidation number changes.

***Step 4.*** Multiply one or both of these numbers by appropriate **factors** to make the electrons lost equal the electron gained, and use the factor as balancing coefficients.

***Step 5.*** Complete the **balancing by inspection**.

Let us have an example.

**Example 1.** Balance the following redox reactions using oxidation number method.

A.) $Cu\ (s) + AgCl(aq) \rightarrow CuCl_2\ (aq) + Ag(s)$

B.) $PbS\ (s) + O_2\ (g) \rightarrow PbO(s) + SO_2(g)$

## Example 1.A

## Balancing Redox Reaction Using Oxidation Number Method

**Balance the redox reaction below using Oxidation Number Method:**

$$Cu\ (s) + AgCl(aq) \rightarrow CuCl_2\ (aq) + Ag(s)$$

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|---|---|
| Step 1. Assign oxidation numbers in every element of the reaction. | Oxidation numbers:<br>$\overset{0}{Cu}\ (s) + \overset{+1\ -1}{AgCl}(aq) \rightarrow \overset{+2\ -1}{CuCl_2}\ (aq) + \overset{0}{Ag}(s)$ |
| Step 2. Identify the oxidized and reduced species based on the change in oxidation numbers.<br><br>Step 3. Compute electron lost and electron gained.<br><br>Note: You can opt to draw lines between the atoms for you to track the changes in the number of gained or lost electrons. | Oxidation numbers:<br>$\overset{0}{Cu}\ (s) + \overset{+1\ -1}{AgCl}(aq) \rightarrow \overset{+2\ -1}{CuCl_2}\ (aq) + \overset{0}{Ag}(s)$<br>Lost 2e-<br>Gained 1e-<br><br>The oxidation number of copper (Cu) increased from 0 (Cu Metal) to +2 (in $CuCl_2$). This means that Cu was oxidized. On the other hand, the oxidation number of silver (Ag) decreased from +1 (in AgCl) to 0 (in Ag). This means that AgCl was |

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| | reduced. Note that Cl does not change its charge (-1) which means that Cl from AgCl and $CuCl_2$ is a spectator ion.<br><br>**Spectator ions** are ions that are part of any reactant solution, but they do not participate in any reactions. This means that these spectator ions can be seen in both product and reactant in any given chemical reaction |
| Step 4. Multiply by factors to make electron gained equals to the electron lost. And use the factors as coefficients. | $\boldsymbol{Cu\,(s) + \; 2AgCl(aq) \rightarrow CuCl_2\,(aq) + 2Ag(s)}$<br><br>Copper lost 2e-, so the 1 e- gained by silver should be multiplied by 2. We then put the resulting factor as a coefficient in AgCl and Ag |
| Step 5. Complete the reaction by inspection. | $\boldsymbol{Cu\,(s) + \; 2AgCl(aq) \rightarrow CuCl_2\,(aq) + 2Ag(s)}$<br><br>In this case since the reaction is already balanced we do not need to balance the reaction by inspection. |

## Example 1.B

## Balancing Redox Reaction Using Oxidation Number Method

**Balance the redox reaction below using Oxidation Number Method:**

$$PbS\,(s) + O_2\,(g) \rightarrow PbO(s) + SO_2(g)$$

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| Step 1. Assign oxidation numbers in every element of the reaction: | Oxidation numbers:<br>+2 -2 0 +2 -2 +4 -2<br>$\boldsymbol{PbS\,(s) + O_2\,(g) \rightarrow PbO(s) + SO_2(g)}$ |

| | |
|---|---|
| Step 2. Identify the oxidized and reduced species based on the change in oxidation numbers. | Oxidation numbers:<br>+2 -2 0 +2 -2 +4-2<br>$\boldsymbol{PbS}\,(s) + \boldsymbol{O_2}\,(g) \rightarrow \boldsymbol{PbO}(s) + \boldsymbol{SO_2}(g)$<br>Gained 4e- (2e- per O)<br>Lost 6 e- |
| Step 3. Compute electron lost and electron gained. Then, draw lines between the atoms. | The oxidation number of S from PbS increased from -2 to +4 in $SO_2$. This means PbS lost 6 electrons thus PbS undergoes oxidation. On the other hand, oxygen from its free gaseous state ($O_2$) decreased its oxidation state from 0 to -2 in either of the two forms PbO and $SO_2$. This means that $O_2$ gained 2 electrons thus it undergoes reduction.<br><br>Sulfur lost 6 electrons. Each oxygen gained 2 electrons making a total of 4 electrons gained |
| Step 4. Multiply by factors to make electron gained equals to the electron lost. And use the factors as coefficients. | $\boldsymbol{PbS}\,(s) + \frac{3}{2}\boldsymbol{O_2}\,(g) \rightarrow \boldsymbol{PbO}(s) + \boldsymbol{SO_2}(g)$<br><br>The S atom loses 6e-, and each O in $O_2$ gains 2e-. , for a total of 4e- . Thus placing a 3/2 coefficient before $O_2$ gives 3 atoms of oxygen with 2 e- each for a total of 6e-. |
| Step 5. Complete the reaction by inspection. | $\boldsymbol{2PbS}\,(s) + \boldsymbol{3O_2}\,(g) \rightarrow \boldsymbol{2PbO}(s) + \boldsymbol{2SO_2}(g)$<br><br>All the atoms are already balanced. However, we have a coefficient that is a fraction. To eliminate the coefficient that is fraction we need to multiply the whole chemical reaction by two. |

The next method in balancing redox reaction is the **ion-electron method** or the **half-reaction method.** The ion-electron method is used when the reaction is usually expressed with ions in aqueous solutions in either in acidic or basic conditions.

**These are the steps in balancing using the half-reaction method. (Tro, 2017)**

***Step 1.*** Assign oxidation states to all atoms and determine the substance that is reduced or oxidized.

***Step 2.*** Separate the two half-reactions (reduction and oxidation).

***Step 3.*** Balance the following half-reactions in the following reactions with respect to mass in the following order

- Balance all elements other H and O
- Balance O by adding $H_2O$ on the side with the least number of oxygen atom.
- Balance H by adding $H^+$ on the side with the least number of hydrogen atom.
- <u>In basic medium</u>, neutralize $H^+$ by adding $OH^-$ on both sides of the reaction with the same amount of $H^+$ ion. (Remember when $H^+$ and $OH^-$ combines it becomes $H_2O$ and cancel $H_2O$ that can be seen on both sides of the equation)

***Step 4.*** Balance each half reaction with respect to charge by adding electrons on the side which has a more positive charge to balance out the charges.

***Step 5.*** Make the number of electrons in both half-reaction equal by multiplying on and both half reaction with a whole number.

***Step 6.*** Add the two half-reactions together canceling electrons and other substances that are present in both sides of the reaction.

***Step 7.*** To check whether your reaction is balance, verify whether the elements and charge on both sides of the reaction is equal.

**Example 2.**

Balance the following reactions using half-reaction method.

A.) $MnO_4^-(aq) + C_2O_4^{2-}\ (aq) \rightarrow\ Mn^{2+}(aq) + CO_2(aq)$ (in acidic medium)

B.) $I^-(aq) + MnO_4^-(aq) \rightarrow\ I_2\ (aq) + MnO_2(s)$(in basic medium)

| **Example 2.A** | **Balancing Redox Reaction Using Half-reaction Method** |
|---|---|
| **Balance the redox reaction below using half-reaction method:** $MnO_4^-(aq) + C_2O_4^{2-}(aq) \rightarrow Mn^{2+}(aq) + CO_2(aq)$ (in acidic medium) | |
| Step 1. Assign oxidation numbers in every element of the reaction and determine the substance that is reduced or oxidized. | Oxidation numbers: $\overset{+7\ -2}{MnO_4^-}(aq) + \overset{+3\ -2}{C_2O_4^{2-}}(aq) \rightarrow \overset{+2}{Mn^{2+}}(aq) + \overset{+4\ -2}{CO_2}(aq)$<br>Substance reduced: $MnO_4^-$ (Mn gained 5e-)<br>Substance oxidized: $C_2O_4^{2-}$ (C lost 1e-) |
| Step 2. Separate the two half-reactions (reduction and oxidation). | Oxidation half-reaction:<br>$C_2O_4^{2-}(aq) \rightarrow CO_2(aq)$<br>Reduction half-reaction:<br>$MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$ |
| Step 3. Balance the following half-reactions in the following reactions with respect to mass in the following order<br>➢ Balance all elements other H and O | $C_2O_4^{2-}(aq) \rightarrow 2CO_2(aq)$<br>$MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$ |

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| ➢ Balance O by adding $H_2O$ on the side with the least number of oxygen atom. | $C_2O_4^{2-}\,(aq) \rightarrow 2CO_2(aq)$<br>$MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l)$<br><br>To balance oxygen we added 4 $H_2O$ to product side of the second half-reaction to compensate for the 4 excess oxygen (O) that is found in permanganate ion ($MnO_4^-$)<br><br>$C_2O_4^{2-}\,(aq) \rightarrow 2CO_2(aq)$<br>$8H^+ + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l)$ |
| ➢ Balance H by adding $H^+$ on the side with the least number of hydrogen atom. | To balance hydrogen we added 8 $H^+$ ion to the reactant side of the second half-reaction to compensate the excess hydrogen atom brought about by the added 4 $H_2O$. |
| Step 4. Balance each half reaction with respect to charge by adding electrons on the side which has a more positive charge to balance out the charges. | <u>For oxidation half-reaction:</u><br><br>$C_2O_4^{2-}\,(aq) \rightarrow 2CO_2(aq)$<br>Charges: -2 2(0)<br><br>Reactants Products<br>Total Charges: -2 0<br><br>Since the charges are not balanced, the reactant side (-2) and product side (0), we add 2 electrons to the side which is more positive side (products) to have a balance in charges.<br><br>$C_2O_4^{2-}\,(aq) \rightarrow 2CO_2(aq) + 2e^-$<br>Charges: -2 2(0) -2<br><br>Reactants Products<br>Total Charges: -2 -2<br><br>By adding the appropriate number of electrons to the side which is more positive, we can balance out the charges of the reaction. To check whether you are on the right track always remember that in an oxidation reaction electrons should be seen on the product side of the equation. (Remember LEORA)<br><br><u>For reduction half-reaction :</u><br><br>$8H^+ + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l)$<br>Charges: 8(+1) -1 +2 4(0)<br><br>Reactants Products<br>Total Charges: +7 +2 |

| | |
|---|---|
| | Since the charges are not balanced, the reactant side (+7) and product side (+2), we add 5 electrons to the side which is more positive (reactant side) to have a balance in charges.<br><br>$$5e^- + 8H^+ + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l)$$<br>Charges: -5 8(+1) -1 +2 4(0)<br><br>Reactants Products<br>Total Charges: +2 +2<br><br>By adding the appropriate number of electrons to the side which is more positive, we can balance out the charges of the reaction. To check whether you are on the right track, always remember that in a reduction reaction electrons should be seen on the reactant side of the equation. (Remember GEROA) |
| Step 5. Make the number of electrons in both half-reaction equal by multiplying on and both half reaction with a whole number. | .<br><br>$$(C_2O_4^{2-}(aq) \rightarrow 2CO_2(aq) + 2e^-)5$$<br>$$(5e^- + 8H^+ + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l))2$$<br><br>To balance the number of electron lost and gained with multiply the 1st half-reaction by 5 and the second half reaction by 2 so that the corresponding number of electrons found in every half-reaction is equal to 10. |
| Step 6. Add the two half-reactions together canceling electrons and other substances that are present in both sides of the reaction. | $$5C_2O_4^{2-}(aq) \rightarrow 10CO_2(aq) + \cancel{10e^-}$$<br>$$\cancel{10e^-} + 16H^+ + 2MnO_4^-(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l)$$<br><br>$$5C_2O_4^{2-}(aq) + 16H^+ + 2MnO_4^-(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 10CO_2(aq)$$<br><br>To check whether the charge of the whole reaction is balanced, we just multiply the coefficients of the ions or compounds with their charges in the reactant side and product side. |

| Ion/ Compound in the reactant | Charge | Ion/ Compound in the Product | Charge |
|---|---|---|---|
| $C_2O_4^{2-}$ | 5(-2) | $Mn^{2+}$ | 2(+2) |
| $H^+$ | 16(+1) | $H_2O$ | 8(0) |
| $MnO_4^-$ | 2(-1) | $CO_2$ | 10(0) |
| Total charge: | +4 | Total charge: | +4 |

Step 7. To check whether your reaction is balance, verify whether the elements and charge on both sides of the reaction is equal.

| Reactants | Products |
|---|---|
| 10 C | 10 C |
| 28 O | 28 O |
| 16 H | 16 H |
| 2 Mn | 2 Mn |
| Charge (+4) | Charge (+4) |

## Example 2.B Balancing Redox Reaction Using Half-reaction Method

**Balance the redox reaction below using half-reaction method:**

$$I^-(aq) + MnO_4^-(aq) \rightarrow I_2\,(aq) + MnO_2(s) \quad \text{(in basic medium)}$$

Step 1. Assign oxidation numbers in every element of the reaction and determine the substance that is reduced or oxidized.

Oxidation numbers:

$$\overset{-1}{\boldsymbol{I^-}}\boldsymbol{(aq)} + \boldsymbol{\overset{+7}{Mn}\overset{-2}{O_4^-}(aq)} \rightarrow \boldsymbol{\overset{0}{I_2}\,(aq)} + \boldsymbol{\overset{+4}{Mn}\overset{-2}{O_2}(s)}$$

Substance reduced: $MnO_4^-$ (Mn gained 3e-)

Substance oxidized: $I^-$ (C lost 1e-)

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|---|---|
| Step 2. Separate the two half-reactions (reduction and oxidation). | Oxidation half-reaction:<br>$I^-(aq) \rightarrow I_2\ (aq)$<br><br>Reduction half-reaction:<br>$MnO_4^-(aq) \rightarrow MnO_2\ (s)$ |
| Step 3. Balance the following half-reactions in the following reactions with respect to mass in the following order<br>➢ Balance all elements other H and O | $2I^-(aq) \rightarrow I_2\ (aq)$<br>$MnO_4^-(aq) \rightarrow MnO_2\ (s)$ |
| ➢ Balance O by adding $H_2O$ on the side with the least number of oxygen atom. | $2I^-(aq) \rightarrow I_2\ (aq)$<br>$MnO_4^-(aq) \rightarrow MnO_2\ (s) + 2H_2O\ (l)$<br><br>To balance oxygen we added 2 $H_2O$ to product side of the second half-reaction to compensate for the 2 excess oxygen (O) that is found in permanganate ion ($MnO_4^-$) |
| ➢ Balance H by adding $H^+$ on the side with the least number of hydrogen atom. | $2I^-(aq) \rightarrow I_2\ (aq)$<br>$4H^+ + MnO_4^-(aq) \rightarrow MnO_2(aq) + 2H_2O(l)$<br><br>To balance hydrogen we added 4 $H^+$ ion to the reactant side of the second half-reaction to compensate the excess hydrogen atom brought about by the added 2 $H_2O$. |
| ➢ <u>In basic medium</u>, neutralize $H^+$ by adding $OH^-$ on both sides of | $2I^-(aq) \rightarrow I_2\ (aq)$<br>$4OH^- + 4H^+ + MnO_4^-(aq)$<br>$\rightarrow MnO_2(aq) + 2H_2O(l) + 4OH^-$ |

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<td>the reaction with the same amount of $H^+$ ion. (Remember when $H^+$ and $OH^-$ combines it becomes $H_2O$). Then we cancel common water molecules on both side</td>
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$$2I^-(aq) \rightarrow I_2\,(aq)$$

$$\underset{2}{\cancel{4}}H_2O\,(l) + MnO_4^-(aq) \rightarrow MnO_2(aq) + \cancel{2H_2O(l)} + 4OH^-$$

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<td>Step 4. Balance each half reaction with respect to charge by adding electrons on the side which has a more positive charge to balance out the charges.</td>
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<u>For oxidation half-reaction:</u>

$$2I^-(aq) \rightarrow I_2\,(aq)$$

Charges: 2(-1) 0

| | Reactants | Products |
|---|---|---|
| Total Charges: | -2 | 0 |

Since the charges are not balanced, the reactant side (-2) and product side (0), we add 2 electrons to the side which is more positive side (product) to have a balance in charges.

$$2I^-(aq) \rightarrow I_2\,(aq) + 2e^-$$

Charges: 2(-1) 0 -2

| | Reactants | Products |
|---|---|---|
| Total Charges: | -2 | -2 |

By adding the appropriate number of electrons to the side which is more positive, we can balance out the charges of the reaction. To check whether you are on the right track always remember that in an oxidation reaction electrons should be seen on the product side of the equation. (Remember LEORA)

<u>For reduction half-reaction:</u>

$$2H_2O\,(l) + MnO_4^-(aq) \rightarrow MnO_2(aq) + 4OH^-$$

Charges: 2(0) -1 0 4(-1)

| | Reactants | Products |
|---|---|---|
| Total Charges: | -1 | -4 |

Since the charges are not balanced, the

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<td>reactant side (-1) and product side (-4), we add 3 electrons to the side which is more positive side (reactant) to have a balance in charges.<br><br>$3e^- + 2H_2O\ (l) + MnO_4^-(aq) \rightarrow MnO_2(aq) + 4OH^-$<br><br>By adding the appropriate number of electrons to the side which is more positive, we can balance out the charges of the reaction. To check whether you are on the right track always remember that in a reduction reaction electrons should be seen on the product side of the equation. (Remember GEROA)</td>
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<td>Step 5. Make the number of electrons in both half-reaction equal by multiplying on and both half reaction with a whole number.</td>
<td>.<br><br>$(\ 2I^-(aq) \rightarrow I_2\ (aq) + 2e^-)3$<br>$2(3e^- + 2H_2O\ (l) + MnO_4^-(aq) \rightarrow MnO_2(aq) + 4OH^-)$</td>
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<td>Step 6. Add the two half-reactions together canceling electrons and other substances that are present in both sides of the reaction.</td>
<td>$6I^-(aq) \rightarrow 3I_2\ (aq) + \cancel{6e^-}$<br>$\cancel{6e^-} + 4H_2O\ (l) + 2MnO_4^-(aq) \rightarrow 2\ MnO_2(aq) + 8OH^-)$<br><br>$6I^-(aq) + 4H_2O\ (l) + 2MnO_4^-(aq)$<br>$\rightarrow 3I_2\ (aq) + 2MnO_2(aq) + 8OH^-$<br><br>To check whether the charge of the whole reaction is balanced, we just multiply the coefficients of the ions or compounds with their charges in the reactant side and product side.<br><br>
<table>
<tr><th>Ion/ Compound in the reactant</th><th>Charge</th><th>Ion/ Compound in the Product</th><th>Charge</th></tr>
<tr><td>$I^-$</td><td>6(-1)</td><td>$I_2$</td><td>3(0)</td></tr>
<tr><td>$H_2O$</td><td>4(0)</td><td>$MnO_2$</td><td>2(0)</td></tr>
<tr><td>$MnO_4^-$</td><td>2(-1)</td><td>$OH^-$</td><td>8(-1)</td></tr>
<tr><td>Total charge:</td><td>-8</td><td>Total charge:</td><td>-8</td></tr>
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Step 7. To check whether your reaction is balance, verify whether the elements and charge on both sides of the reaction is equal.

| Reactants | Products |
|---|---|
| 6 I | 6 I |
| 12 O | 12 O |
| 8 H | 8 H |
| 2 Mn | 2 Mn |
| Charge (   8) | Charge (-8) |

Things to remember in balancing using half-reaction method:

- If you do not know the oxidation state of an element, you can solve it by initially assuming its value as X.
  Ex. To solve for the oxidation number of $MnO_4^-$
  Let X = oxidation number of Mn
  (X)+ 4(Oxidation state of O) = (total charge of ion or compound)
  So, X + 4(-2) = -1
  X = +7 ; the oxidation state of Mn in $MnO_4^-$ is (+7)
- To check whether you are on the right track remember our mnemonics for redox reaction. "LEORA and GEROA". This means that during a reduction process the substance gains an electron, which is why in reduction electron, should be seen in the reactant. On the other hand during oxidation, electron is released so, electron in an oxidation reaction should be seen in the product side.

**Navigate**

Work on the following exercises to find out if you understood the lesson. Problems 1A, 2A, and 2B are NON-GRADED. While problems 1B, 1C, 2C, and 2D are GRADED and to be submitted on the platform provided by your teacher.

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1.) Balance the following redox reactions using oxidation state method:
   A.) $Ni(s) + HCl(s) \rightarrow NiCl_2(s) + H_2(s)$
   B.) $Cu\ (s) + HNO_3(aq) \rightarrow Cu(NO_3)_2\ (aq) + NO_2(g) + \ H_2O\ (g)$
   C.) $Fe_2O_3\ (s) + CO\ (g) \rightarrow Fe(s) + CO_2(g)$

2.) Balance the following redox reaction using half-reaction method
   A.) $Cu(s) + NO_3^-(aq) \rightarrow \ Cu^{+2}(aq) + NO_2(g)$ in acidic medium
   B.) $CN^-(aq) + MnO_4^-(aq) \rightarrow CNO^-(aq) + MnO_2(s)$in basic medium
   C.) $Cr_2O_7^{2-}(aq) + \ Cl^-(aq) \rightarrow Cr^{3+}(aq) + Cl_2(g)$in acidic medium
   D.) $ClO^-(aq) + Cr(OH)_4^-\ (aq) \rightarrow Cr\ O_4^{2-}\ (aq) + \ Cl^-\ (aq)$in basic medium

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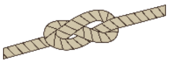
| **Knot**<br> | Here are some of the significant key ideas that you should remember about redox reaction.<br>▪ Like any other chemical reactions balancing redox reaction is important to account all the atoms present in a chemical reaction, since atoms cannot be created nor destroyed in a chemical reaction.<br>▪ There are two methods of balancing redox reactions. We have oxidation number method and half-reaction method.<br>▪ The important aspect of balancing redox reactions is that not only the mass or the amount of atoms should be balanced, the charge of the participant (substance reduced and oxidized) should also be balanced as well. | 2 | |

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|---|---|
| Prepared by: DEXTER VILE J. LAURDEN | Reviewed by: PRINCESS ANN DIGNENENG |
| Position: Special Science Teacher II | Position: Special Science Teacher II |
| Campus: CALABARZON | Campus: CLC |