BALANCING REDOX REACTIONS

Oxidation Number Method

OXIDATION NUMBER METHOD

Redox reactions can be balanced using:

a) Half Reactions

b) Oxidation numbers

Goal: **To balance the electrons that are gained and lost**

NOTE: When half-reactions are <u>not</u> provided, and a reaction does <u>not</u> occur in either acidic or basic solution, then the oxidation number method can be used to balance a redox reaction

OXIDATION NUMBER METHOD

Example #1

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution (recall the half-reaction method rules)

$$\underset{0}{Al_{(s)}} + \underset{+4\ \ -2}{MnO_{2(s)}} \rightarrow \underset{+3\ \ -2}{Al_2O_{3(s)}} + \underset{0}{Mn_{(s)}}$$

OXIDATION NUMBER METHOD

Example #1

1. Assign oxidation numbers to each element.
2. Identify the <u>increase</u> and <u>decrease</u> in oxidation numbers.
3. Determine the <u>smallest whole-number ratio</u> of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to <u>balance the oxidized and reduced reactants</u>.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in <u>acidic</u> or <u>basic</u> solution

+3

$$\underset{0}{Al_{(s)}} + \underset{+4\;\;-2}{MnO_{2(s)}} \rightarrow \underset{+3\;\;-2}{Al_2O_{3(s)}} + \underset{0}{Mn_{(s)}}$$

-4

OXIDATION NUMBER METHOD

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5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\overbrace{\underset{0}{Al_{(s)}} + \underset{+4\ \ -2}{MnO_{2(s)}} \rightarrow \underset{+3\ \ -2}{Al_2O_{3(s)}}}^{+3} + \underset{0}{Mn_{(s)}}$$

$$\underbrace{MnO_{2(s)} \rightarrow Al_2O_{3(s)} + Mn_{(s)}}_{-4}$$

Ratio = 4:3

Each Al atom loses 3, .: 4 Al loses 12
Each Mn atom gains 4, .: 3 Mn gains 12

OXIDATION NUMBER METHOD

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5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

+3

$$4Al_{(s)} + 3MnO_{2(s)} \rightarrow Al_2O_{3(s)} + Mn_{(s)}$$

$$\underset{0}{Al} \quad \underset{+4}{Mn}\ \underset{-2}{O} \quad \underset{+3}{Al}\ \underset{-2}{O} \quad \underset{0}{Mn}$$

-4

Ratio = 4:3

OXIDATION NUMBER METHOD

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6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\overbrace{4Al_{(s)} + 3MnO_{2(s)} \rightarrow 2Al_2O_{3(s)}}^{+3} + 3Mn_{(s)}$$

Oxidation numbers: Al 0; Mn +4, O -2; Al +3, O -2; Mn 0

$$3\underbrace{MnO_{2(s)} \rightarrow 2Al_2O_{3(s)} + 3Mn_{(s)}}_{-4}$$

Ratio = 4:3

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5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$4Al_{(s)} + 3MnO_{2(s)} \rightarrow 2Al_2O_{3(s)} + 3Mn_{(s)}$$

OXIDATION NUMBER METHOD

Example #2

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\underset{-3\ +1}{NH_{3(g)}} + \underset{0}{O_{2(g)}} \rightarrow \underset{+4\ -2}{NO_{2(g)}} + \underset{+1\ \ -2}{H_2O_{(l)}}$$

OXIDATION NUMBER METHOD

Example #2

1. Assign oxidation numbers to each element.
2. Identify the <u>increase</u> and <u>decrease</u> in oxidation numbers.
3. Determine the <u>smallest whole-number ratio</u> of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to <u>balance the oxidized and reduced reactants</u>.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in <u>acidic</u> or <u>basic</u> solution

+7

$$\underset{-3\ +1}{NH_{3(g)}} + \underset{0}{O_{2(g)}} \rightarrow \underset{+4\ -2}{NO_{2(g)}} + \underset{+1\ \ -2}{H_2O_{(l)}}$$

-2 x 2

$$\underset{0}{O_2} \rightarrow \underset{-2}{O}$$

Each O gains 2 electrons, for a total of 4 electrons

OXIDATION NUMBER METHOD

Example #2

1. Assign oxidation numbers to each element.
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4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\overbrace{\underset{-3\ +1}{NH_{3(g)}} + \underset{0}{O_{2(g)}} \rightarrow \underset{+4\ -2}{NO_{2(g)}}}^{+7} + \underset{+1\ \ -2}{H_2O_{(l)}}$$

Change for O ($O_2 \rightarrow H_2O$): -4

$$\underset{0}{O} \rightarrow \underset{-2}{O_2}$$

If it was the other way around, though, then you don't need to multiply the change in electrons by 2

OXIDATION NUMBER METHOD

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5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\overset{+7}{\overbrace{\underset{-3\ +1}{NH_{3(g)}} + \underset{0}{O_{2(g)}} \rightarrow \underset{+4\ -2}{NO_{2(g)}}}} + \underset{+1\ \ -2}{H_2O_{(l)}}$$

-4 (change from O_2: 0 to -2)

Ratio = 4:7

Each N atom loses 7, .: 4 N loses 28
Each O atom gains 2 → each O_2 gains 4, .: 7 O_2 gains 28

OXIDATION NUMBER METHOD

Example #2

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\overbrace{4\underset{-3\ +1}{NH_{3(g)}} + 7\underset{0}{O_{2(g)}} \rightarrow \underset{+4\ -2}{NO_{2(g)}}}^{+7} + \underset{+1\ \ -2}{H_2O_{(l)}}$$

$$O_2 \rightarrow H_2O: -4$$

Ratio = 4:7

Each N atom loses 7, .: 4 N loses 28

Each O atom gains 2 → each O_2 gains 4, .: 7 O_2 gains 28

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4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\overbrace{4\,\underset{-3\;+1}{NH_{3(g)}} + 7\,\underset{0}{O_{2(g)}} \rightarrow 4\,\underset{+4\;-2}{NO_{2(g)}}}^{+7} + 6\,\underset{+1\;\;-2}{H_2O_{(l)}}$$

-4

Ratio = 4:7

Each N atom loses 7, .: 4 N loses 28

Each O atom gains 2 → each O_2 gains 4, .: 7 O_2 gains 28

OXIDATION NUMBER METHOD

Example #2

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$4NH_{3(g)} + 7O_{2(g)} \rightarrow 4NO_{2(g)} + 6H_2O_{(l)}$$

OXIDATION NUMBER METHOD

Example #3: In acidic solution

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\underset{0}{Cu_{(s)}} + \underset{+5\ -2}{NO_3^-{}_{(aq)}} \rightarrow \underset{+4\ -2}{NO_{2(g)}} + \underset{+2}{Cu^{2+}{}_{(aq)}}$$

OXIDATION NUMBER METHOD

Example #3: In acidic solution

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\overbrace{Cu_{(s)} + NO_3^-{}_{(aq)} \rightarrow NO_{2(g)} + Cu^{2+}{}_{(aq)}}^{+2}$$

$$\underset{0}{Cu_{(s)}} + \underset{+5\ -2}{NO_3^-{}_{(aq)}} \rightarrow \underset{+4\ -2}{NO_{2(g)}} + \underset{+2}{Cu^{2+}{}_{(aq)}}$$

$$\underbrace{+5 \rightarrow +4}_{-1}$$

OXIDATION NUMBER METHOD

Example #3: In acidic solution

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\overset{0}{Cu_{(s)}} + \overset{+5\ -2}{NO_3^-{}_{(aq)}} \rightarrow \overset{+4\ -2}{NO_{2(g)}} + \overset{+2}{Cu^{2+}{}_{(aq)}}$$

Cu: +2

N: -1

Ratio = 1:2

Each Cu atom gains 2, .: 1 Cu gains 2

Each N atom loses 1, .: 2 N loses 2

OXIDATION NUMBER METHOD

Example #3: In acidic solution

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

+2

$$\underset{0}{Cu_{(s)}} + 2\,\underset{+5\ -2}{NO_3^-{}_{(aq)}} \rightarrow \underset{+4\ -2}{NO_{2(g)}} + \underset{+2}{Cu^{2+}{}_{(aq)}}$$

-1

Ratio = 1:2

Each Cu atom gains 2, .: 1 Cu gains 2

Each N atom loses 1, .: 2 N loses 2

OXIDATION NUMBER METHOD

Example #3: In acidic solution

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\overset{+2}{\overbrace{\underset{0}{Cu_{(s)}} + 2\,\underbrace{\underset{+5\ -2}{NO_3^-{}_{(aq)}} \rightarrow 2\,\underset{+4\ -2}{NO_{2(g)}}}_{-1} + \underset{+2}{Cu^{2+}{}_{(aq)}}}}$$

Ratio = 1:2

Each Cu atom gains 2, .: 1 Cu gains 2
Each N atom loses 1, .: 2 N loses 2

OXIDATION NUMBER METHOD

Example #3: In acidic solution

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
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5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$4H^+ + Cu_{(s)} + 2NO_3^-{}_{(aq)} \rightarrow 2NO_{2(g)} + Cu^{2+}{}_{(aq)} + 2H_2O$$

1) If ions are present, balance the charge by adding H^+ for acid solutions or OH^- for basic solutions.
2) Finish balancing the equation by adding H_2O.

OXIDATION NUMBER METHOD

Example #3: In acidic solution

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4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$4H^+ + Cu_{(s)} + 2NO_3^-{}_{(aq)} \rightarrow 2NO_{2(g)} + Cu^{2+}{}_{(aq)} + 2H_2O$$

OXIDATION NUMBER METHOD

Example #4: Fractions

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$\underset{-2\ \ +1\ \ -2\ \ +1}{C_3H_7OH} + \underset{+1\ \ +6\ \ -2}{Na_2Cr_2O_7} + \underset{+1\ \ +6\ \ -2}{H_2SO_4} \rightarrow \underset{+3\ \ +6\ \ -2}{Cr_2(SO_4)_3} + \underset{+1\ \ +6\ \ -2}{Na_2SO_4} + \underset{+1\ \ -2}{H_2O} + \underset{+1\ \ -\frac{2}{3}\ \ +1\ \ -2}{HC_3H_5O_2}$$

OXIDATION NUMBER METHOD

Example #4: Fractions

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$-3 \times 2 = \mathbf{6}$$

$$C_3H_7OH + Na_2Cr_2O_7 + H_2SO_4 \rightarrow Cr_2(SO_4)_3 + Na_2SO_4 + H_2O + HC_3H_5O_2$$

C_3H_7OH	$Na_2Cr_2O_7$	H_2SO_4	$Cr_2(SO_4)_3$	Na_2SO_4	H_2O	$HC_3H_5O_2$
-2 +1 -2 +1	+1 +6 -2	+1 +6 -2	+3 +6 -2	+1 +6 -2	+1 -2	+1 $-\frac{2}{3}$ +1 -2

$$+\frac{4}{3} \times 3 = \mathbf{4}$$

$$-\frac{2}{3} - (-2) = +\frac{4}{3}$$

OXIDATION NUMBER METHOD

Example #4: Fractions

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5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

6

$$C_3H_7OH + Na_2Cr_2O_7 + H_2SO_4 \rightarrow Cr_2(SO_4)_3 + Na_2SO_4 + H_2O + HC_3H_5O_2$$

$$\begin{array}{ccccccc} -2\;\;+1\;-2\;+1 & +1\;\;\;+6\;\;-2 & +1\;\;+6\;-2 & +3\;\;\;+6\;-2 & +1\;\;+6\;-2 & +1\;\;-2 & +1\;-\frac{2}{3}\;\;+1\;\;-2 \end{array}$$

4

Ratio = 6:4

= 3:2

OXIDATION NUMBER METHOD

Example #4: Fractions

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
4. Use the determined ratio from step 3 to balance the oxidized and reduced reactants.
5. Balance the other elements/compounds by inspection.
6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

6

$$\mathbf{3}\ C_3H_7OH + \mathbf{2}Na_2Cr_2O_7 + H_2SO_4 \rightarrow Cr_2(SO_4)_3 + Na_2SO_4 + H_2O + HC_3H_5O_2$$

-2 +1-2 +1   +1 +6 -2   +1 +6-2   +3 +6-2   +1 +6-2   +1 -2   +1 $-\frac{2}{3}$ +1 -2

4

Ratio = 6:4

= **3:2**

OXIDATION NUMBER METHOD

Example #4: Fractions

1. Assign oxidation numbers to each element.
2. Identify the increase and decrease in oxidation numbers.
3. Determine the smallest whole-number ratio of the oxidized and reduced elements so that the total increase in oxidation numbers equals the total decrease in oxidation numbers.
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6

$$\mathbf{3}\,\underset{-2\ \ +1\ -2\ +1}{C_3H_7OH} + \mathbf{2}\,\underset{+1\ \ +6\ -2}{Na_2Cr_2O_7} + \mathbf{8}\,\underset{+1\ -6\ -2}{H_2SO_4} \rightarrow \mathbf{2}\,\underset{+3\ \ -6\ -2}{Cr_2(SO_4)_3} + \mathbf{2}\,\underset{+1\ -6\ -2}{Na_2SO_4} + \mathbf{11}\,\underset{+1\ -2}{H_2O} + \mathbf{3}\,\underset{+1\,-\frac{2}{3}\ \ +1\ -2}{HC_3H_5O_2}$$

4

Ratio = 6:4

= **3:2**

OXIDATION NUMBER METHOD

Example #4: Fractions

1. Assign oxidation numbers to each element.
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6. Add H_2O, H^+, and OH^- as needed to balance reactions in acidic or basic solution

$$3\,C_3H_7OH + 2\,Na_2Cr_2O_7 + 8\,H_2SO_4 \rightarrow 2\,Cr_2(SO_4)_3 + 2\,Na_2SO_4 + 11\,H_2O + 3\,HC_3H_5O_2$$