Chemistry 1

Volume 4

Worksheet 20

Balancing Redox Reactions in Acidic Solution

Ion/Electron Method

Part 5

Rules for balancing redox reactions in acidic solution:

1. Write the net ionic equation.
2. Write half-reactions.
3. Balance atoms other than hydrogen and oxygen.
4. Balance oxygen by adding H_2O.
5. Balance H by adding H^+.
6. Balance the charges by adding e^-.
    a. Add up the total charge on each side.
    b. Add electrons to the more positive side to balance the charge on both sides.
7. Multiply to balance the electrons gained/lost between the two half-reactions and add reactions together.
8. Cancel common species on both sides of the reaction.

1. Balance the redox reaction below that is in acidic solution:

$$NO + H_5IO_6 \rightarrow NO_3^- + IO_3^-$$

2. Balance the redox reaction below that is in acidic solution:

$$Cr_2O_7^{2-} + C_2H_6O \rightarrow Cr^{3+} + C_2H_4O$$

3. Balance the redox reaction below that is in acidic solution:

$$CH_3CHO + MnO_4^- \rightarrow Mn^{2+} + CH_3COOH$$

4. Balance the redox reaction below that is in acidic solution:

$$O^{2-} + F_2 \rightarrow O_2 + F^-$$

5. Balance the redox reaction below that is in acidic solution:

$$PbO_2 + I_2 \rightarrow Pb^{2+} + IO_3^-$$

Answer Key

1. Balance the redox reaction below that is in acidic solution:

$$NO + H_5IO_6 \rightarrow NO_3^- + IO_3^-$$

Step 1:
The net ionic equation is already written, so we can go to rule 2.

Step 2:
Break into half-reactions.

$NO \rightarrow NO_3^-$

$H_5IO_6 \rightarrow IO_3^-$

Step 3:
All elements other than O or H are already balanced.

$NO \rightarrow NO_3^-$

$H_5IO_6 \rightarrow IO_3^-$

Step 4:
Balance O by adding H_2O.

$2\ H_2O + NO \rightarrow NO_3^-$

$H_5IO_6 \rightarrow IO_3^- + 3\ H_2O$

Step 5:
Balance H by adding H^+.

$2\ H_2O + NO \rightarrow NO_3^- + 4\ H^+$

$H^+ + H_5IO_6 \rightarrow IO_3^- + 3\ H_2O$

Step 6:
Balance charges by adding e^-.

$2\ H_2O + NO \rightarrow NO_3^- + 4\ H^+ + 3\ e^-$

$2\ e^- + H^+ + H_5IO_6 \rightarrow IO_3^- + 3\ H_2O$

Step 7:
Multiply the half-reactions in order to balance electrons and add them together.

$2 \times [\ 2\ H_2O + NO \rightarrow NO_3^- + 4\ H^+ + 3\ e^-\ ]$

$3 \times [\ 2\ e^- + H^+ + H_5IO_6 \rightarrow IO_3^- + 3\ H_2O\ ]$

$6\ e^- + 3\ H^+ + 3\ H_5IO_6 + 4\ H_2O + 2\ NO \rightarrow 2\ NO_3^- + 8\ H^+ + 6\ e^- + 3\ IO_3^- + 9\ H_2O$

Step 8:
Cancel common species.

~~$6\ e^-$~~ $+$ ~~$3\ H^+$~~ $+ 3\ H_5IO_6 +$ ~~$4\ H_2O$~~ $+ 2\ NO \rightarrow 2\ NO_3^- + 5$ ~~8~~ $H^+ +$ ~~$6\ e^-$~~ $+ 3\ IO_3^- + 5$ ~~9~~ H_2O

Now is a good time to make sure all elements and charges are balanced.

$3\ H_5IO_6 + 2\ NO \rightarrow 2\ NO_3^- + 5\ H^+ + 3\ IO_3^- + 5\ H_2O$

Reactants		Products	
H	15	H	15
I	3	I	3
O	20	O	20
N	2	N	2
Charge	0	Charge	0

Everything is balanced!

Correct answer:

$3\ H_5IO_6 + 2\ NO \rightarrow 2\ NO_3^- + 5\ H^+ + 3\ IO_3^- + 5\ H_2O$

2. Balance the redox reaction below that is in acidic solution:

$$Cr_2O_7^{2-} + C_2H_6O \rightarrow Cr^{3+} + C_2H_4O$$

Step 1:
The net ionic equation is already written, so we can go to rule 2.

Step 2:
Break into half-reactions.

$C_2H_6O \rightarrow C_2H_4O$

$Cr_2O_7^{2-} \rightarrow Cr^{3+}$

Step 3:
Balance all elements other than O or H.

$C_2H_6O \rightarrow C_2H_4O$

$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+}$

Step 4:
Balance O by adding H_2O.

$C_2H_6O \rightarrow C_2H_4O$

$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

Step 5:
Balance H by adding H^+.

$C_2H_6O \rightarrow C_2H_4O + 2\ H^+$

$14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

Step 6:
Balance charges by adding e^-.

$C_2H_6O \rightarrow C_2H_4O + 2\ H^+ + 2\ e^-$

$6\ e^- + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

Step 7:
Multiply the half-reactions in order to balance electrons and add them together.

3 x [$C_2H_6O \rightarrow C_2H_4O + 2\ H^+ + 2\ e^-$]

$6\ e^- + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

$6\ e^- + 14\ H^+ + Cr_2O_7^{2-} + 3\ C_2H_6O \rightarrow 3\ C_2H_4O + 6\ H^+ + 6\ e^- + 2\ Cr^{3+} + 7\ H_2O$

Step 8:
Cancel common species.

~~$6\ e^-$~~ + 8 ~~14~~ $H^+ + Cr_2O_7^{2-} + 3\ C_2H_6O \rightarrow 3\ C_2H_4O +$ ~~$6\ H^+$~~ + ~~$6\ e^-$~~ $+ 2\ Cr^{3+} + 7\ H_2O$

Make sure all elements and charges are balanced.

$8\ H^+ + Cr_2O_7^{2-} + 3\ C_2H_6O \rightarrow 3\ C_2H_4O + 2\ Cr^{3+} + 7\ H_2O$

Reactants		Products	
H	26	H	26
Cr	2	Cr	2
O	10	O	10
C	6	C	6
Charge	+6	Charge	+6

Everything is balanced!

Correct answer:

$8\ H^+ + Cr_2O_7^{2-} + 3\ C_2H_6O \rightarrow 3\ C_2H_4O + 2\ Cr^{3+} + 7\ H_2O$

3. Balance the redox reaction below that is in acidic solution:

$$CH_3CHO + MnO_4^- \rightarrow Mn^{2+} + CH_3COOH$$

Step 1:
The net ionic equation is already written, so we can go to rule 2.

Step 2:
Break into half-reactions.

$CH_3CHO \rightarrow CH_3COOH$

$MnO_4^- \rightarrow Mn^{2+}$

Step 3:
All elements other than O or H are already balanced.

$CH_3CHO \rightarrow CH_3COOH$

$MnO_4^- \rightarrow Mn^{2+}$

Step 4:
Balance O by adding H_2O.

$H_2O + CH_3CHO \rightarrow CH_3COOH$

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

Step 5:
Balance H by adding H^+.

$H_2O + CH_3CHO \rightarrow CH_3COOH + 2\ H^+$

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

Step 6:
Balance charges by adding e^-.

$H_2O + CH_3CHO \rightarrow CH_3COOH + 2\ H^+ + 2\ e^-$

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

Step 7:
Multiply the half-reactions in order to balance electrons and add them together.

5 x [$H_2O + CH_3CHO \rightarrow CH_3COOH + 2\ H^+ + 2\ e^-$]

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

$10\ e^- + 16\ H^+ + 2\ MnO_4^- + 5\ H_2O + 5\ CH_3CHO \rightarrow 5\ CH_3COOH + 10\ H^+ + 10\ e^- + 2\ Mn^{2+} + 8\ H_2O$

Step 8:
Cancel common species.

~~$10\ e^-$~~ + 6 ~~16~~ H^+ + 2 MnO_4^- + ~~5 H_2O~~ + 5 $CH_3CHO \rightarrow$ 5 CH_3COOH + ~~10 H^+~~ + ~~10 e^-~~ + 2 Mn^{2+} + 3 ~~8~~ H_2O

Make sure all elements and charges are balanced.

$6\ H^+ + 2\ MnO_4^- + 5\ CH_3CHO \rightarrow 5\ CH_3COOH + 2\ Mn^{2+} + 3\ H_2O$

Reactants		Products	
H	26	H	26
Mn	2	Mn	2
O	13	O	13
C	10	C	10
Charge	+4	Charge	+4

Everything is balanced!

Correct answer:

$6\ H^+ + 2\ MnO_4^- + 5\ CH_3CHO \rightarrow 5\ CH_3COOH + 2\ Mn^{2+} + 3\ H_2O$

4. Balance the redox reaction below that is in acidic solution:

$$O^{2-} + F_2 \rightarrow O_2 + F^-$$

Step 1:
The net ionic equation is already written, so we can go to rule 2.

Step 2:
Break into half-reactions.

$O^{2-} \rightarrow O_2$

$F_2 \rightarrow F^-$

Step 3:
Balance all elements other than O or H.

$O^{2-} \rightarrow O_2$

$F_2 \rightarrow 2\ F^-$

Step 4:
Balance O by adding H_2O.

$H_2O + O^{2-} \rightarrow O_2$

$F_2 \rightarrow 2\ F^-$

Step 5:
Balance H by adding H^+.

$H_2O + O^{2-} \rightarrow O_2 + 2\ H^+$

$F_2 \rightarrow 2\ F^-$

Step 6:
Balance charges by adding e^-.

$H_2O + O^{2-} \rightarrow O_2 + 2\ H^+ + 4\ e^-$

$2\ e^- + F_2 \rightarrow 2\ F^-$

Step 7:
Multiply the half-reactions in order to balance electrons and add them together.

$H_2O + O^{2-} \rightarrow O_2 + 2\ H^+ + 4\ e^-$

$2 \times [\ 2\ e^- + F_2 \rightarrow 2\ F^-\ ]$

$4\ e^- + 2\ F_2 + H_2O + O^{2-} \rightarrow O_2 + 2\ H^+ + 4\ e^- + 4\ F^-$

Step 8:
Cancel common species.

~~$4\ e^-$~~ $+ 2\ F_2 + H_2O + O^{2-} \rightarrow O_2 + 2\ H^+ +$ ~~$4\ e^-$~~ $+ 4\ F^-$

Make sure all elements and charges are balanced.

$2\ F_2 + H_2O + O^{2-} \rightarrow O_2 + 2\ H^+ + 4\ F^-$

Reactants		Products	
F	4	F	4
H	2	H	2
O	2	O	2
Charge	-2	Charge	-2

Everything is balanced!

Correct answer:

$2\ F_2 + H_2O + O^{2-} \rightarrow O_2 + 2\ H^+ + 4\ F^-$

5. Balance the redox reaction below that is in acidic solution:

$$PbO_2 + I_2 \rightarrow Pb^{2+} + IO_3^-$$

Step 1:
The net ionic equation is already written, so we can go to rule 2.

Step 2:
Break into half-reactions.

$PbO_2 \rightarrow Pb^{2+}$

$I_2 \rightarrow IO_3^-$

Step 3:
Balance all elements other than O or H.

$PbO_2 \rightarrow Pb^{2+}$

$I_2 \rightarrow 2\ IO_3^-$

Step 4:
Balance O by adding H_2O.

$PbO_2 \rightarrow Pb^{2+} + 2\ H_2O$

$6\ H_2O + I_2 \rightarrow 2\ IO_3^-$

Step 5:
Balance H by adding H^+.

$4\ H^+ + PbO_2 \rightarrow Pb^{2+} + 2\ H_2O$

$6\ H_2O + I_2 \rightarrow 2\ IO_3^- + 12\ H^+$

Step 6:
Balance charges by adding e^-.

$2\ e^- + 4\ H^+ + PbO_2 \rightarrow Pb^{2+} + 2\ H_2O$

$6\ H_2O + I_2 \rightarrow 2\ IO_3^- + 12\ H^+ + 10\ e^-$

Step 7:
Multiply the half-reactions in order to balance electrons and add them together.

5 x [2 e^- + 4 H^+ + PbO_2 → Pb^{2+} + 2 H_2O]

6 H_2O + I_2 → 2 IO_3^- + 12 H^+ + 10 e^-

6 H_2O + I_2 + 10 e^- + 20 H^+ + 5 PbO_2 → 5 Pb^{2+} + 10 H_2O + 2 IO_3^- + 12 H^+ + 10 e^-

Step 8:
Cancel common species.

~~6 H_2O~~ + I_2 + ~~10 e^-~~ + 8 ~~20~~ H^+ + 5 PbO_2 → 5 Pb^{2+} + 4 ~~10~~ H_2O + 2 IO_3^- + ~~12 H^+~~ + ~~10 e^-~~

Make sure all elements and charges are balanced.

I_2 + 8 H^+ + 5 PbO_2 → 5 Pb^{2+} + 4 H_2O + 2 IO_3^-

Reactants		Products	
I	2	I	2
H	8	H	8
Pb	5	Pb	5
O	10	O	10
Charge	+8	Charge	+8

Everything is balanced!

Correct answer:

I_2 + 8 H^+ + 5 PbO_2 → 5 Pb^{2+} + 4 H_2O + 2 IO_3^-