4/19/2019

Redox rxns.

- ox. states (chem-1) ~review

how do we balance redox eqs? → almost always in acidic/basic environment.

Procedure...

Method

Example

ox state ±#

1. Determine ox state for each atom.

(+2) (+7)(-2) (+3) (+2) charge #±

$Fe^{2+}_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + Mn^{2+}_{(aq)}$

(in acid)

(+2) (+3) ox # ↑ when ox!

2. Separated into ½-rxns (ox-red)

$Fe^{2+} \xrightarrow{ox} Fe^{3+}$

(+7) (+2) ox # ↓ when red!

$MnO_4^- \xrightarrow{red} Mn^{2+}$

3. Balance each ½-rxn by:

a) Coefficients for everything except O, H

b) O by adding H_2O

c) H by adding H^+

(ox) $Fe^{2+} \rightarrow Fe^{3+}$ ✓

(red) $MnO_4^- \rightarrow Mn^{2+}$ coeff ✓

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

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

4. Balance charge by adding e^-s to more +ve side

(ox) $Fe^{2+} \rightarrow Fe^{3+} + e^-$ ✓

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

5. × ½-rxns by small whole #'s so that we have same # e^-

$5\times\left(\quad Fe^{2+} \longrightarrow Fe^{3+} + e^- \quad\right)$

$1\times\left(5e^- + 8H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O\right)$

⊕

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6. Add ½ rxns + cancel out

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

7. Check it's balanced!

8. If we are balancing under basic conditions, add OH^- to both sides, to neut H^+: $H^+ + OH^- \rightarrow H_2O$

$+8OH^-$ $\qquad$ $+8OH^-$

$\cancel{8}H_2O$

4

final eq: $5Fe^{2+}_{(aq)} + 4H_2O_{(l)} + MnO_4^-{}_{(aq)} \xrightarrow{5e^-} 5Fe^{3+}_{(aq)} + Mn^{2+}_{(aq)} + 8OH^-_{(aq)}$

Example 19.1 Half-Reaction Method of Balancing Aqueous Redox Equations in Acidic Solution

Balance the redox equation:

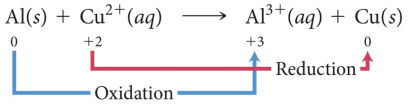
$$Al(s) + Cu^{2+}(aq) \longrightarrow Al^{3+}(aq) + Cu(s)$$

Procedure For…

Half-Reaction Method of Balancing Aqueous Redox Equations in Acidic Solution

General Procedure

Step 1 *Assign oxidation states* to all atoms and identify the substances being oxidized and reduced.



Step 2 *Separate the overall reaction into two half-reactions*: one for oxidation and one for reduction.

Oxidation: $Al(s) \longrightarrow Al^{3+}(aq)$
Reduction: $Cu^{2+}(aq) \longrightarrow Cu(s)$

Example 19.1 Half-Reaction Method of Balancing Aqueous Redox Equations in Acidic Solution

Continued

Step 3 *Balance each half-reaction with respect to mass* in the following order:
- Balance all elements other than H and O.
- Balance O by adding H_2O.
- Balance H by adding H^+.

All elements are balanced, so proceed to the next step.

Step 4 *Balance each half-reaction with respect to charge* by adding electrons. (Make the sum of the charges on both sides of the equation equal by adding as many electrons as necessary.)

$$Al(s) \longrightarrow Al^{3+}(aq) + 3\ e^-$$
$$2\ e^- + Cu^{2+}(aq) \longrightarrow Cu(s)$$

Step 5 *Make the number of electrons in both half-reactions equal* by multiplying one or both half-reactions by a small whole number.

$$2[Al(s) \longrightarrow Al3+(aq) + 3\ e^-]$$
$$2\ Al(s) \longrightarrow 2\ Al3+(aq) + 6\ e^-$$
$$3[2\ e^- + Cu^{2+}(aq) \longrightarrow Cu(s)]$$
$$6\ e^- + 3\ Cu^{2+}(aq) \longrightarrow 3\ Cu(s)$$

Example 19.1 Half-Reaction Method of Balancing Aqueous Redox Equations in Acidic Solution

Continued

Step 6 *Add the two half-reactions together*, canceling electrons and other species as necessary.

$$2\,Al(s) \longrightarrow 2\,Al^{3+}(aq) + 6\,e^-$$
$$6\,e^- + 3\,Cu^{2+}(aq) \longrightarrow 3\,Cu(s)$$
$$2\,Al(s) + 3\,Cu^{2+}(aq) \longrightarrow 2\,Al^{3+}(aq) + 3\,Cu(s)$$

Step 7 *Verify that the reaction is balanced* with respect to both mass and charge.

Reactants	Products
2 Al	2 Al
3 Cu	3 Cu
+6 charge	+6 charge

For Practice 19.1

Balance the redox reaction in acidic solution:

$$H^+(aq) + Cr(s) \longrightarrow H_2(g) + Cr^{2+}(aq)$$

Example 19.2 Half-Reaction Method of Balancing Aqueous Redox Equations in Acidic Solution

Balance the redox equation:

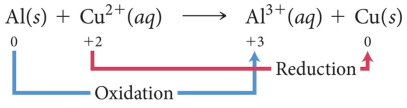
$$Fe^{2+}(aq) + MnO_4^-(aq) \longrightarrow Fe^{3+}(aq) + Mn^{2+}(aq)$$

Procedure For…

Half-Reaction Method of Balancing Aqueous Redox Equations in Acidic Solution

General Procedure

Step 1 *Assign oxidation states* to all atoms and identify the substances being oxidized and reduced.



Step 2 *Separate the overall reaction into two half-reactions*: one for oxidation and one for reduction.

Oxidation: $Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq)$
Reduction: $MnO_4^-(aq) \longrightarrow Mn^{2+}(aq)$

Example 19.2 Half-Reaction Method of Balancing Aqueous Redox Equations in Acidic Solution

Continued

Step 3 *Balance each half-reaction with respect to mass* in the following order:

- Balance all elements other than H and O.
- Balance O by adding H_2O.
- Balance H by adding H^+.

All elements other than H and O are balanced, so proceed to balance H and O.

$$Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq)$$
$$MnO_4^-(aq) \longrightarrow Mn^{2+}(aq) + 4\ H_2O(l)$$
$$8\ H^+(aq) + MnO_4^-(aq) \longrightarrow Mn^{2+}(aq) + 4\ H_2O(l)$$

Step 4 *Balance each half-reaction with respect to charge* by adding electrons. (Make the sum of the charges on both sides of the equation equal by adding as many electrons as necessary.)

$$Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + 1\ e^-$$
$$5\ e^- + 8\ H^+(aq) + MnO_4^-(aq) \longrightarrow Mn^{2+}(aq) + 4\ H_2O(l)$$

Example 19.2 Half-Reaction Method of Balancing Aqueous Redox Equations in Acidic Solution

Continued

Step 5 *Make the number of electrons in both half-reactions equal* by multiplying one or both half-reactions by a small whole number.

$$5[Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + 1\ e^-]$$
$$5\ Fe^{2+}(aq) \longrightarrow 5\ Fe3+(aq) + 5\ e^-$$
$$5\ e^- + 8\ H^+(aq) + MnO_4^-(aq) \longrightarrow Mn^{2+}(aq) + 4\ H_2O(l)$$

Step 6 *Add the two half-reactions together*, canceling electrons and other species as necessary.

$$5\ Fe^{2+}(aq) \longrightarrow 5\ Fe^{3+}(aq) + 5\ e^-$$
$$5\ e^- + 8\ H^+(aq) + MnO_4^-(aq) \longrightarrow Mn^{2+}(aq) + 4\ H_2O(l)$$
$$5\ Fe^{2+}(aq) + 8\ H^+(aq) + MnO_4^-(aq) \longrightarrow 5\ Fe^{3+}(aq) + Mn^{2+}(aq) + 4\ H_2O(l)$$

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Example 19.2 Half-Reaction Method of Balancing Aqueous Redox Equations in Acidic Solution

Continued

Step 7 *Verify that the reaction is balanced* with respect to both mass and charge.

Reactants	Products
5 Fe	5 Fe
8 H	8 H
1 Mn	1 Mn
4 O	4 O
+17 charge	+17 charge

For Practice 19.2

Balance the redox reaction in acidic solution:

$$Cu(s) + NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + NO_2(g)$$

Example 19.3 Balancing Redox Reactions Occurring in Basic Solution

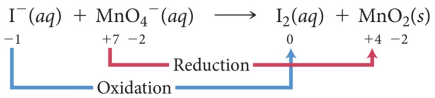
Balance the equation occurring in basic solution:

$$I^-(aq) + MnO_4^-(aq) \longrightarrow I_2(aq) + MnO_2(s)$$

Solution

To balance redox reactions occurring in basic solution, follow the half-reaction method outlined in Examples 19.1 and 19.2, but add an extra step to neutralize the acid with OH- as shown in step 3.

Step 1 Assign oxidation states.



Step 2 Separate the overall reaction into two half-reactions.

Oxidation: $I^-(aq) \longrightarrow I_2(aq)$
Reduction: $MnO_4^-(aq) \longrightarrow MnO_2(s)$

Example 19.3 Balancing Redox Reactions Occurring in Basic Solution

Continued

Step 3 Balance each half-reaction with respect to mass.

- Balance all elements other than H and O.
- Balance O by adding H_2O.
- Balance H by adding H^+.
- Neutralize H^+ by adding enough OH^- to neutralize each H^+. Add the same number of OH^- ions to each side of the equation.

$$\left\{\begin{array}{l} 2\ I^-(aq) \longrightarrow I_2(aq) \\ MnO_4^-(aq) \longrightarrow MnO_2(s) \end{array}\right.$$

$$\left\{\begin{array}{l} 2\ I^-(aq) \longrightarrow I_2(aq) \\ MnO_4^-(aq) \longrightarrow MnO_2(s) + 2\ H_2O(l) \end{array}\right.$$

$$\left\{\begin{array}{l} 2\ I^-(aq) \longrightarrow I_2(aq) \\ 4\ H^+(aq) + MnO_4^-(aq) \longrightarrow MnO_2(s) + 2\ H_2O(l) \end{array}\right.$$

$$\left\{\begin{array}{l} 2\ I^-(aq) \longrightarrow I_2(aq) \\ \underbrace{4\ H^+(aq) + 4\ OH^-(aq)}_{4\ H_2O(l)} + MnO_4^-(aq) \longrightarrow MnO_2(s) + 2\ H_2O(l) + 4\ OH^-(aq) \end{array}\right.$$

Example 19.3 Balancing Redox Reactions Occurring in Basic Solution

Continued

Step 4 Balance each half-reaction with respect to charge.

$$2\ I^-(aq) \longrightarrow I_2(aq) + 2\ e^-$$
$$4\ H_2O(l) + MnO_4^-(aq) + 3\ e^- \longrightarrow MnO_2(s) + 2\ H_2O(l) + 4\ OH^-(aq)$$

Step 5 Make the number of electrons in both half-reactions equal.

$$3[2\ I^-(aq) \longrightarrow I_2(aq) + 2\ e^-]$$
$$6\ I^-(aq) \longrightarrow 3\ I_2(aq) + 6\ e^-$$
$$2[4\ H_2O(l) + MnO_4^-(aq) + 3\ e^- \longrightarrow MnO_2(s) + 2\ H_2O(l) + 4\ OH^-(aq)]$$
$$8\ H_2O(l) + 2\ MnO_4^-(aq) + 6\ e^- \longrightarrow 2\ MnO_2(s) + 4\ H_2O(l) + 8\ OH^-(aq)$$

Step 6 Add the half-reactions together.

$$6\ I^-(aq) \longrightarrow 3\ I_2(aq) + 6\ e^-$$

$$8\ H_2O(l) + 2\ MnO_4^-(aq) + 6\ e^- \longrightarrow 2\ MnO_2(s) + 4\ H_2O(l) + 8\ OH^-(aq)$$

$$6\ I^-(aq) + 4\ H_2O(l) + 2\ MnO_4^-(aq) \longrightarrow 3\ I_2(aq) + 2\ MnO_2(s) + 8\ OH^-(aq)$$

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Example 19.3 Balancing Redox Reactions Occurring in Basic Solution

Continued

Step 7 Verify that the reaction is balanced.

Reactants	Products
6 I	6 I
8 H	8 H
2 Mn	2 Mn
12 O	12 O
−8 charge	−8 charge

For Practice 19.3

Balance the following redox reaction occurring in basic solution:

$$ClO^-(aq) + Cr(OH)_4^-(aq) \longrightarrow CrO_4^{2-}(aq) + Cl^-(aq)$$

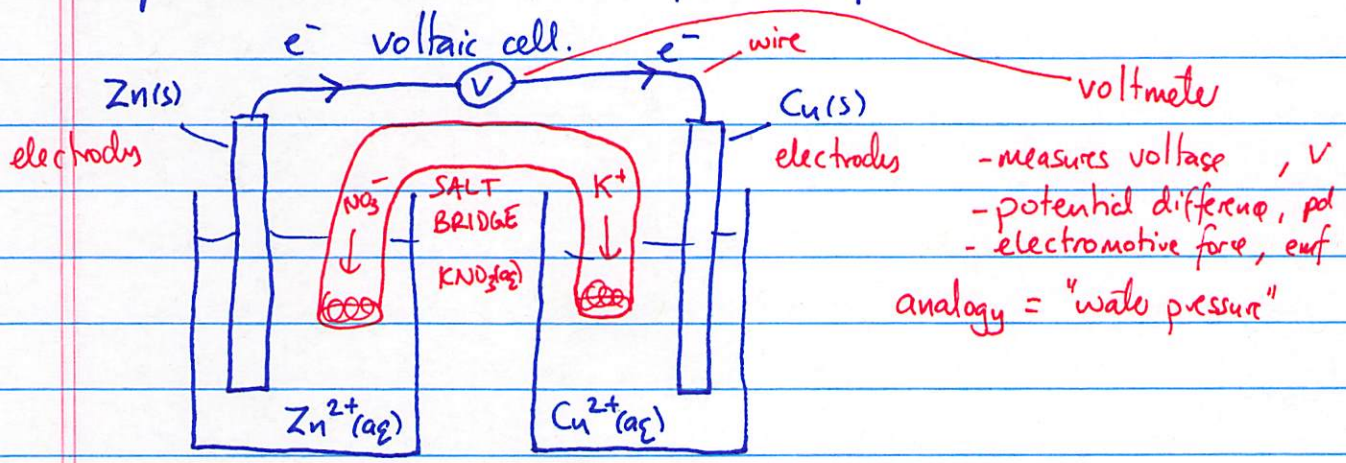
Voltaic (Galvanic) Cells

- generate elec. from spont. chem rxns.

ex: $\overset{(0)}{Zn(s)} + \overset{(+2)}{Cu^{2+}(aq)} \xrightarrow{2e^-} \overset{(+2)}{Zn^{2+}(aq)} + \overset{(0)}{Cu(s)}$

ox red

if we want to harness e⁻'s, must separate reactants!



$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ OXIDATION ANODE

$2e^- + Cu^{2+}(aq) \rightarrow Cu(s)$ REDUCTION CATHODE

RC-COCA

salt bridge keeps sol^{n}s elec. neutral.