مكتبة الفكر

Chapter 4

Redox Reactions

Revision Paper

SECTION 1: OXIDATION AND REDUCTION
SECTION 2: BALANCING REDOX EQUATIONS

Oxidation and Reduction

Chapter 4
Lesson 1 – Revision Paper

Section Summary

- Oxidation and reduction are complementary – as a substance is oxidized, another substance is reduced.

- Oxidation-reduction reactions involve the transfer of electrons from one atom to another.

- When an atom or ion is reduced, its oxidation number decreases. When an atom or ion is oxidized, its oxidation number increases.

- In oxidation-reduction reactions involving molecular compounds (and polyatomic ions with covalent bonds), the **more-electronegative atoms** are treated as if they are **reduced**. The **less-electronegative atoms** are treated as if they are **oxidized**.

Figure 1

01

Classify the reaction between magnesium and oxygen.

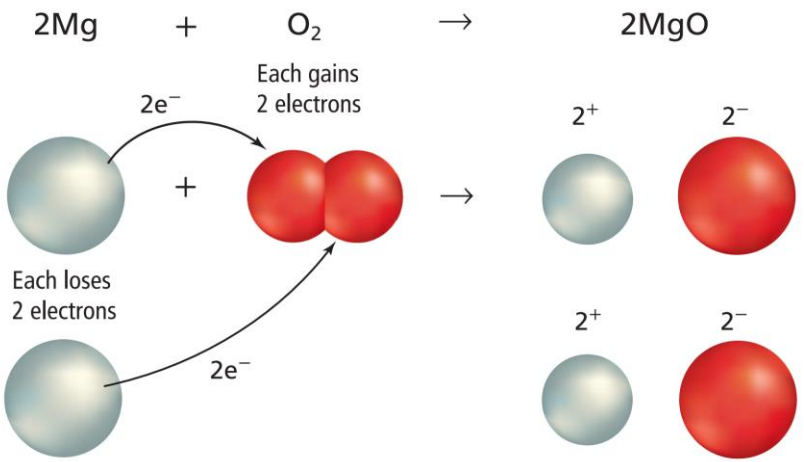
a) Synthesis

b) Decomposition

c) Combustion

d) Single-replacement

e) Double-replacement



Page 134

4.1 Oxidation and Reduction

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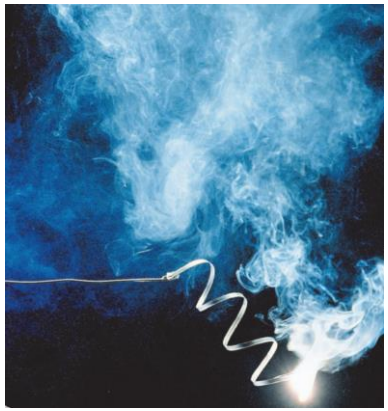

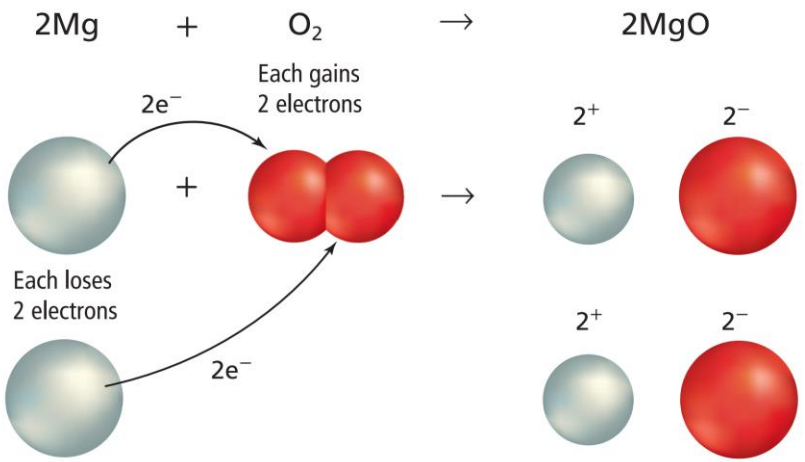

Page 134

4.1 Oxidation and Reduction

Reading Check

02

Determine which element is more likely to gain electrons, potassium or chlorine?

a) Potassium

b) Chlorine

c) Both have the same chance

d) Cannot be determined

Reading Check

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Determine which element is more likely to gain electrons, potassium or chlorine?

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Figure 4

Predict which element would be the strongest oxidizing agent.

a) Li

b) Ba

c) F2

d) I2

e) Cs

f) O2

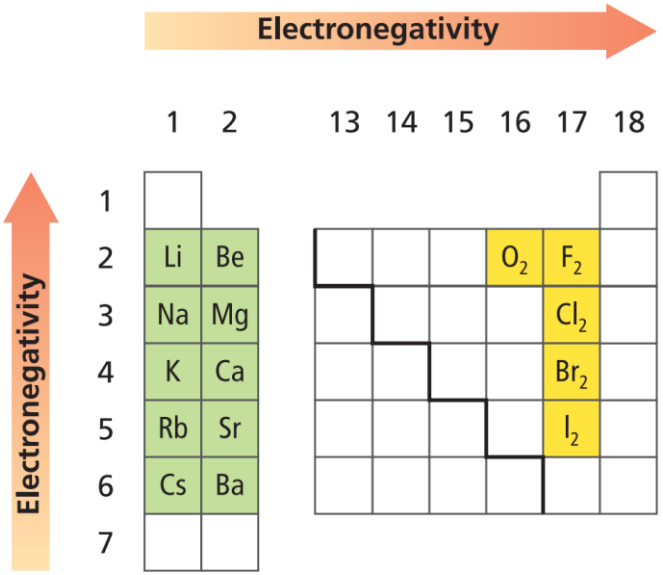
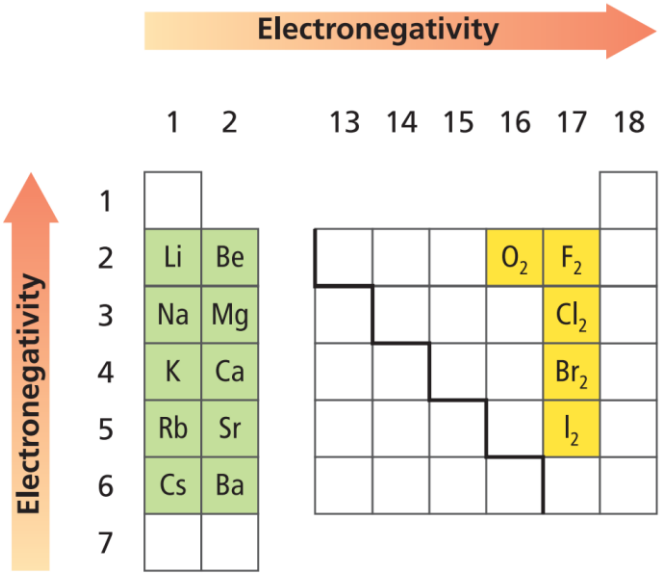


Figure 4

Predict which element would be the strongest oxidizing agent.



a) Li

b) Ba

c) F2

d) I2

e) Cs

f) O2

Page 138

4.1 Oxidation and Reduction

Figure 4

04

Which is the strongest reducing agent?

a) Li

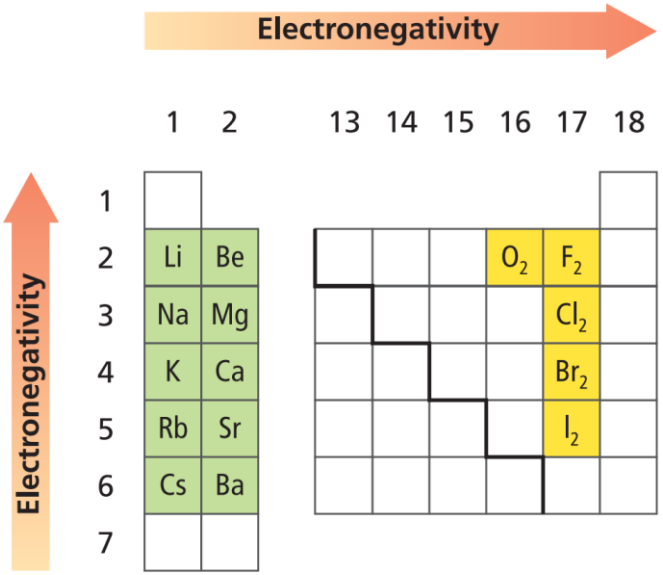
b) Ba

c) Be

d) I2

e) Cs

f) O2



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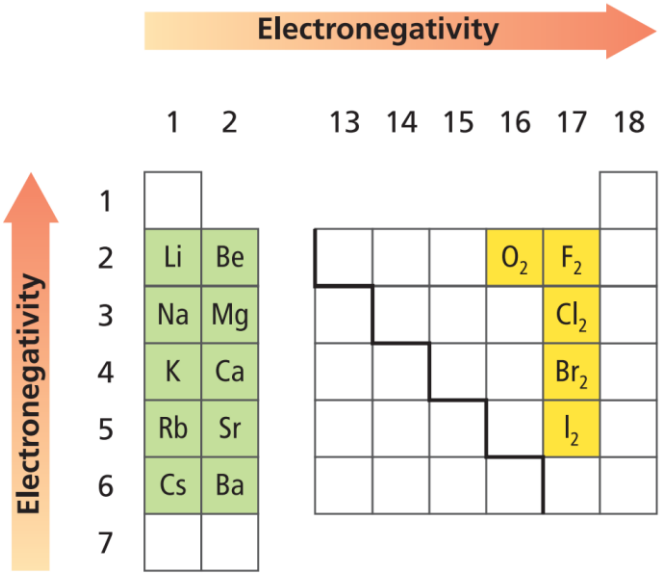

Page 138

4.1 Oxidation and Reduction

Example 1

The following equation represents the redox reaction of aluminum and iron. Identify what is oxidized and what is reduced in this reaction. Identify the oxidizing agent and the reducing agent.

$$2Al + 2Fe^{3+} + 3O^{2-} \longrightarrow 2Fe + 2Al^{3+} + 3O^{2-}$$

Example 1

The following equation represents the redox reaction of aluminum and iron. Identify what is oxidized and what is reduced in this reaction. Identify the oxidizing agent and the reducing agent.

$$2Al + 2Fe^{3+} + 3O^{2-} \rightarrow 2Fe + 2Al^{3+} + 3O^{2-}$$

Identify the oxidation process and the reduction process.

$Al \rightarrow Al^{3+} + 3e^-$ (loss of e^- is oxidation) **The aluminum atom loses three electrons and becomes an aluminum ion.**

$Fe^{3+} + 3e^- \rightarrow Fe$ (gain of e^- is reduction) **The iron ion accepts the three electrons lost from aluminum and becomes an iron atom.**

Al is oxidized and is therefore the reducing agent. Fe^{3+} is reduced and is therefore the oxidizing agent.

Applications 1

Identify each of the following changes as either oxidation or reduction. Recall that **e**$^-$ is the symbol for an electron.

a. $I_2 + 2e^- \longrightarrow 2I^-$

b. $K \longrightarrow K^+ + e^-$

c. $Fe^{2+} \longrightarrow Fe^{3+} + e^-$

d. $Ag^+ + e^- \longrightarrow Ag$

Applications 1

Identify each of the following changes as either oxidation or reduction. Recall that **e**$^-$ is the symbol for an electron.

a. $I_2 + 2e^- \longrightarrow 2I^-$

reduction

b. $K \longrightarrow K^+ + e^-$

oxidation

c. $Fe^{2+} \longrightarrow Fe^{3+} + e^-$

oxidation

d. $Ag^+ + e^- \longrightarrow Ag$

reduction

Applications 2

Identify what is oxidized and what is reduced in the following processes.

a. $2Br^- + Cl_2 \longrightarrow Br_2 + 2Cl^-$

b. $2Ce + 3Cu^{2+} \longrightarrow 3Cu + 2Ce^{3+}$

c. $2Zn + O_2 \longrightarrow 2ZnO$

d. $2Na + 2H^+ \longrightarrow 2Na^+ + H_2$

Page 139

4.1 Oxidation and Reduction

Applications 2

Identify what is oxidized and what is reduced in the following processes.

a. $2Br^- + Cl_2 \longrightarrow Br_2 + 2Cl^-$

Br is oxidized, Cl is reduced

b. $2Ce + 3Cu^{2+} \longrightarrow 3Cu + 2Ce^{3+}$

Ce is oxidized, Cu^{2+} is reduced

c. $2Zn + O_2 \longrightarrow 2ZnO$

Zn is oxidized, O_2 is reduced

d. $2Na + 2H^+ \longrightarrow 2Na^+ + H_2$

H^+ is reduced, Na is oxidized

Page 139

4.1 Oxidation and Reduction

Applications 3

Identify the oxidizing agent and the reducing agent in the following equation. Explain your answer.

$$Fe(s) + Ag^{+}(aq) \longrightarrow Fe^{2+}(aq) + Ag(s)$$

Applications 3

Identify the oxidizing agent and the reducing agent in the following equation. Explain your answer.

$$Fe(s) + Ag^{+}(aq) \longrightarrow Fe^{2+}(aq) + Ag(s)$$

- **Ag^{+}** is **reduced**, **Ag^{+}** is the **oxidizing agent**.
- **Fe** is **oxidized**, **Fe** is the **reducing agent**.

يُعدّ Ag^{+} العامل المؤكسِد، في حين يُعدّ Fe العامل المُختَزِل.
لذا، تُختَزَل أيونات Ag^{+}، وتتأكسد ذرات Fe.

Applications 4

CHALLENGE Identify the oxidizing agent and the reducing agent in each reaction.

a. $Mg + I_2 \longrightarrow MgI_2$

b. $H_2S + Cl_2 \longrightarrow S + 2HCl$

Applications 4

CHALLENGE Identify the oxidizing agent and the reducing agent in each reaction.

a. $Mg + I_2 \longrightarrow MgI_2$

I_2 is the oxidizing agent, Mg is the reducing agent

b. $H_2S + Cl_2 \longrightarrow S + 2HCl$

Cl_2 is the oxidizing agent, H_2S is the reducing agent

Example 2

Use the rules for determining oxidation numbers to find the oxidation number of each element in **potassium chlorate** (**$KClO_3$**) and in a **sulfite ion** (**SO_3^{2-}**).

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Use the rules for determining oxidation numbers to find the oxidation number of each element in **potassium chlorate** (**$KClO_3$**) and in a **sulfite ion** (**SO_3^{2-}**).

Assign the known oxidation numbers to their elements, set the sum of all oxidation numbers to zero or to the ion charge, and solve for the unknown oxidation number.

$(n_K) + (n_{Cl}) + 3\,(n_O) = 0$
$(+1) + (n_{Cl}) + 3(-2) = 0$
$1 + n_{Cl} + (-6) = 0$
$n_{Cl} = +5$

The sum of the oxidation numbers in a neutral compound is zero. For group 1 metals, $n_{element} = +1$. Substitute $n_K = +1$, $n_O = -2$.

Solve for n_{Cl}.

$(n_S) + 3\,(n_O) = -2$
$(n_S) + 3(-2) = -2$
$n_S + (-6) = -2$
$n_S = +4$

The sum of the oxidation numbers in a polyatomic ion equals the charge on the ion. Substitute $n_O = -2$.

Solve for n_S.

Applications 5

Determine the oxidation number of the boldface element in the following formulas for compounds.

a. $\mathrm{NaClO_4}$

b. $\mathrm{AlPO_4}$

c. $\mathrm{HNO_2}$

Page 141

4.1 Oxidation and Reduction

Applications 5

Determine the oxidation number of the boldface element in the following formulas for compounds.

a. $\mathrm{NaClO_4}$

+7

b. $\mathrm{AlPO_4}$

+5

c. $\mathrm{HNO_2}$

+3

Page 141

4.1 Oxidation and Reduction

Applications 6

Determine the oxidation number of the boldface element in the following formulas for ions.

a. $\mathbf{N}H_4^+$

b. $\mathbf{As}O_4^{3-}$

c. $\mathbf{Cr}O_4^{2-}$

Page 141

4.1 Oxidation and Reduction

Applications 6

Determine the oxidation number of the boldface element in the following formulas for ions.

a. $\mathbf{NH_4^+}$

−3

b. $\mathbf{AsO_4^{3-}}$

+5

c. $\mathbf{CrO_4^{2-}}$

+6

Page 141

4.1 Oxidation and Reduction

Applications 7

Determine the oxidation number of nitrogen in each of these molecules.

a. NH_3

b. KCN

c. N_2H_4

Applications 7

Determine the oxidation number of nitrogen in each of these molecules.

a. NH_3

−3

b. KCN

−3

c. N_2H_4

- 2

Applications 8

CHALLENGE Determine the net change of oxidation number of each of the elements in these redox equations.

a. $C + O_2 \longrightarrow CO_2$

b. $Cl_2 + ZnI_2 \longrightarrow ZnCl_2 + I_2$

c. $CdO + CO \longrightarrow Cd + CO_2$

Applications 8

CHALLENGE Determine the net change of oxidation number of each of the elements in these redox equations.

a. $C + O_2 \longrightarrow CO_2$

C, + 4; O, −2

b. $Cl_2 + ZnI_2 \longrightarrow ZnCl_2 + I_2$

I, +1; Cl, −1; Zn, no change

c. $CdO + CO \longrightarrow Cd + CO_2$

C, +2; Cd, −2; O, no change

Review 9

Explain why oxidation and reduction must always occur together.

Review 9

Explain why oxidation and reduction must always occur together.

If an atom loses an electron, some other species must gain the electron.

إذا فقدت ذرة أو أيون ما إلكترونًا، فيجب أن تكتسب ذرات أخرى هذا الإلكترون.

Review 10

Describe the roles of oxidizing agents and reducing agents in a redox reaction. How is each changed in the reaction?

Review 10

Describe the roles of oxidizing agents and reducing agents in a redox reaction. How is each changed in the reaction?

- **An oxidizing agent causes another species to be oxidized by gaining the electrons from it.**
- **A reducing agent causes another species to be reduced by losing electrons to that element.**

- **العامل المؤكسد يتسبب في تأكسد نوع آخر عن طريق اكتساب الإلكترونات منه.**
- **العامل المختزل يتسبب في اختزال نوع آخر عن طريق فقدان إلكترونات إليه.**

Review 11

Write the equation for the reaction of **iron metal** with **hydrobromic acid** to form aqueous **iron(III) bromide** and **hydrogen gas**. Determine the change in oxidation number for the element that is reduced and the element that is oxidized.

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Write the equation for the reaction of **iron metal** with **hydrobromic acid** to form aqueous **iron(III) bromide** and **hydrogen gas**. Determine the change in oxidation number for the element that is reduced and the element that is oxidized.

$$2Fe + 6HBr \longrightarrow 2FeBr_3 + 3H2$$

Fe is oxidized, H is reduced.

Review 12

Determine the oxidation number of the boldface element in these compounds.

a. $\mathrm{HNO_3}$

b. $\mathrm{Ca_3N_2}$

c. $\mathrm{Sb_2O_5}$

d. $\mathrm{CuWO_4}$

Page 142

4.1 Oxidation and Reduction

Review 12

Determine the oxidation number of the boldface element in these compounds.

a. $\mathrm{HNO_3}$

+5

b. $\mathrm{Ca_3N_2}$

−3

c. $\mathrm{Sb_2O_5}$

+5

d. $\mathrm{CuWO_4}$

+6

Page 142

4.1 Oxidation and Reduction

Review 13

Determine the oxidation number of the boldface element in these ions.

a. $\mathbf{I}O_4^-$

b. $\mathbf{Mn}O_4^-$

c. $\mathbf{B}_4O_7^{2-}$

d. $\mathbf{N}H_2^-$

Page 142

4.1 Oxidation and Reduction

Review 13

Determine the oxidation number of the boldface element in these ions.

a. $\mathbf{I}O_4^-$

+7

b. $\mathbf{Mn}O_4^-$

+7

c. $\mathbf{B}_4O_7^{2-}$

+3

d. $\mathbf{N}H_2^-$

−3

Page 142

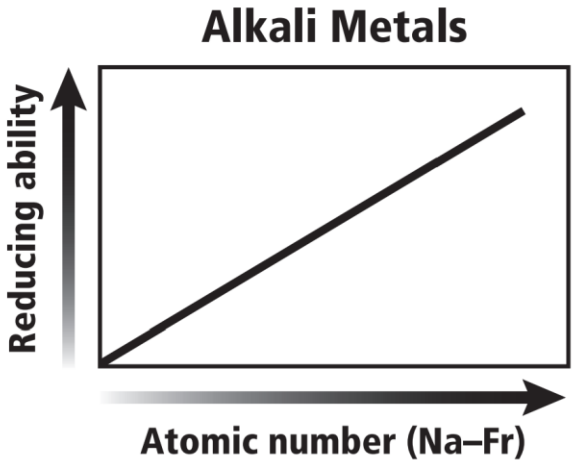
4.1 Oxidation and Reduction

Review 14

Alkali metals are strong reducing agents. Make a graph showing how the reducing abilities of the alkali metals increase or decrease as you move down the family from sodium to francium.

Review 14

Alkali metals are strong reducing agents. Make a graph showing how the reducing abilities of the alkali metals increase or decrease as you move down the family from sodium to francium.



In general, as you move down the periodic table within a family, the tendency to lose electrons increases so the reducing ability increases.

Page 142

4.1 Oxidation and Reduction

Mastering Concepts 33

What is the main characteristic of oxidation-reduction reactions?

a) involve two reactants

b) are decomposition reactions

c) involve the transfer of electrons

d) are synthesis reactions

Page 154

4.1 Oxidation and Reduction

Mastering Concepts 33

What is the main characteristic of oxidation-reduction reactions?

a) involve two reactants

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Page 154

4.1 Oxidation and Reduction

Mastering Concepts 34

Explain why not all oxidation reactions involve oxygen.

Mastering Concepts 34

Explain why not all oxidation reactions involve oxygen.

Originally, the word oxidation referred only to reactions in which a substance combined with oxygen. Today, oxidation is defined as the complete or partial loss of electrons from a reacting substance.

Mastering Concepts 35

In terms of electrons, what happens when an atom is oxidized?

a) Electrons are lost

b) Electrons are gained.

c) Electrons are shared.

d) Electrons are accepted.

Mastering Concepts 35

In terms of electrons, what happens when an atom is oxidized?

a) Electrons are lost.

b) Electrons are gained.

c) Electrons are shared.

d) Electrons are accepted.

Mastering Concepts 35

In terms of electrons, what happens when an atom is reduced?

a) Electrons are donated.

b) Electrons are lost.

c) Electrons are shared.

d) Electrons are gained.

Mastering Concepts 35

In terms of electrons, what happens when an atom is reduced?

a) Electrons are donated.

b) Electrons are lost.

c) Electrons are shared.

d) Electrons are gained.

Mastering Concepts 36

_____ is the number of electrons lost or gained by an atom in an ionic compound when it forms ions

a) Redox number

b) Oxidation number

c) Reduction number

d) Ionic number

Page 154

4.1 Oxidation and Reduction

Mastering Concepts 36

_____ is the number of electrons lost or gained by an atom in an ionic compound when it forms ions

a) Redox number

b) Oxidation number

c) Reduction number

d) Ionic number

Page 154

4.1 Oxidation and Reduction

Mastering Concepts 37

What is the oxidation number of alkaline earth metals in their compounds?

a) +1

b) +2

c) +3

d) -1

Page 154

4.1 Oxidation and Reduction

Mastering Concepts 37

What is the oxidation number of alkaline earth metals in their compounds?

a) +1

b) +2

c) +3

d) -1

Page 154

4.1 Oxidation and Reduction

Mastering Concepts 37

What is the oxidation number of alkali metals in their compounds?

a) -2

b) +1

c) +2

d) +3

Mastering Concepts 37

What is the oxidation number of alkali metals in their compounds?

a) -2

b) +1

c) +2

d) +3

Mastering Concepts 38

How does the oxidation number in an oxidation process relate to the number of electrons lost? How does the change in oxidation number in a reduction process relate to the number of electrons gained?

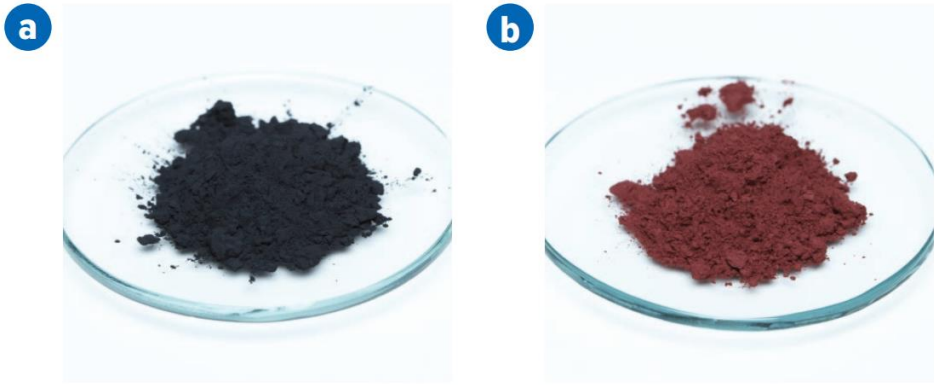
Mastering Concepts 38

How does the oxidation number in an oxidation process relate to the number of electrons lost? How does the change in oxidation number in a reduction process relate to the number of electrons gained?

The change in oxidation number equals the number of electrons lost in oxidation, gained in reduction.

Mastering Concepts 39

What probably accounts for the different forms of copper shown in the figure?



Mastering Concepts 39

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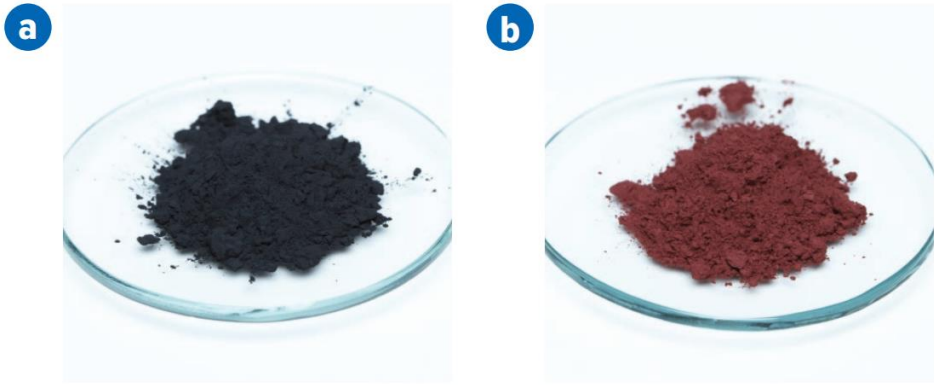


Copper has different oxidation numbers in each form.

Page 154

4.1 Oxidation and Reduction

Mastering Concepts 40

Copper statues, such as the Statue of Liberty, begin to appear green after they have been exposed to air. In this redox process, copper metal reacts with oxygen to form solid copper oxide, which forms the green coating. Write the reaction for this redox process, and identify what is oxidized and what is reduced in the process.

Mastering Concepts 40

Copper statues, such as the Statue of Liberty, begin to appear green after they have been exposed to air. In this redox process, copper metal reacts with oxygen to form solid copper oxide, which forms the green coating. Write the reaction for this redox process, and identify what is oxidized and what is reduced in the process.

$$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$$

Cu is oxidized, O is reduced

Mastering Problems 41

Identify the species oxidized and the species reduced in each of these redox equations.

a. $3Br_2 + 2Ga \longrightarrow 2GaBr_3$

b. $HCl + Zn \longrightarrow ZnCl_2 + H_2$

c. $Mg + N_2 \longrightarrow Mg_3N_2$

Mastering Problems 41

Identify the species oxidized and the species reduced in each of these redox equations.

a. $3Br_2 + 2Ga \longrightarrow 2GaBr_3$

Ga is oxidized, Br_2 is reduced.

b. $HCl + Zn \longrightarrow ZnCl_2 + H_2$

Zn is oxidized, H is reduced.

c. $Mg + N_2 \longrightarrow Mg_3N_2$

Mg is oxidized, N_2 is reduced.

Mastering Problems 42

Identify the oxidizing agent and the reducing agent in each of these redox equations.

a. $N_2 + 3H_2 \longrightarrow 2NH_3$

b. $2Na + I_2 \longrightarrow 2NaI$

Mastering Problems 42

Identify the oxidizing agent and the reducing agent in each of these redox equations.

a. $N_2 + 3H_2 \longrightarrow 2NH_3$

N_2 is the oxidizing agent, H_2 is the reducing agent.

b. $2Na + I_2 \longrightarrow 2NaI$

I_2 is the oxidizing agent, Na is the reducing agent.

Mastering Problems 43

What is the reducing agent in this balanced equation?

$$8H^+ + Sn + 6Cl^- + 4NO_3^- \rightarrow SnCl_6^{2-} + 4NO_2 + 4H_2O$$

Page 154

4.1 Oxidation and Reduction

Mastering Problems 43

What is the reducing agent in this balanced equation?

$$8H^+ + Sn + 6Cl^- + 4NO_3^- \rightarrow SnCl_6^{2-} + 4NO_2 + 4H_2O$$

Sn is the reducing agent.

Mastering Problems 44

What is the oxidation number of manganese in **$KMnO_4$**?

Page 154

4.1 Oxidation and Reduction

Mastering Problems 44

What is the oxidation number of manganese in **$KMnO_4$**?

The oxidation number of manganese in $KMnO_4$ is +7.

Mastering Problems 45

Determine the oxidation number of the boldface element in these substances and ions.

a. $\mathrm{CaCrO_4}$

b. $\mathrm{NaHSO_4}$

c. $\mathrm{NO_2^-}$

d. $\mathrm{BrO_3^-}$

Mastering Problems 45

Determine the oxidation number of the boldface element in these substances and ions.

a. $\mathrm{CaCrO_4}$

+6

b. $\mathrm{NaHSO_4}$

+6

c. $\mathrm{NO_2^-}$

+3

d. $\mathrm{BrO_3^-}$

+5

Page 154

4.1 Oxidation and Reduction

Mastering Problems 46

Identify each of these half-reactions as either oxidation or reduction.

a. $Al \longrightarrow Al^{3+} + 3e^-$

b. $Cu^{2+} + e^- \longrightarrow Cu^+$

Page 154

4.1 Oxidation and Reduction

Mastering Problems 46

Identify each of these half-reactions as either oxidation or reduction.

a. $Al \longrightarrow Al^{3+} + 3e^-$

oxidation

b. $Cu^{2+} + e^- \longrightarrow Cu^+$

reduction

Mastering Problems 47

Which of these equations does not represent a redox reaction? Explain your answer.

a. $LiOH + HNO_3 \longrightarrow LiNO_3 + H_2O$

b. $MgI_2 + Br_2 \longrightarrow MgBr_2 + I_2$

Mastering Problems 47

Which of these equations does not represent a redox reaction? Explain your answer.

a. $LiOH + HNO_3 \longrightarrow LiNO_3 + H_2O$

b. $MgI_2 + Br_2 \longrightarrow MgBr_2 + I_2$

Choice (a) is not redox because none of the atoms in the reaction undergoes a change in oxidation number.

Mastering Problems 48

Determine the oxidation number of nitrogen in each of these molecules or ions.

a. NO_3^-

b. N_2O

c. NF_3

Page 154

4.1 Oxidation and Reduction

Mastering Problems 48

Determine the oxidation number of nitrogen in each of these molecules or ions.

a. NO_3^-

+5

b. N_2O

+1

c. NF_3

+3

Mastering Problems 49

Determine the oxidation number of each element in these compounds or ions.

a. $Au_2(SeO_4)_3$ (gold(III) selenate)

b. $Ni(CN)_2$ (nickel(II) cyanide)

Mastering Problems 49

Determine the oxidation number of each element in these compounds or ions.

a. $Au_2(SeO_4)_3$ (gold(III) selenate)

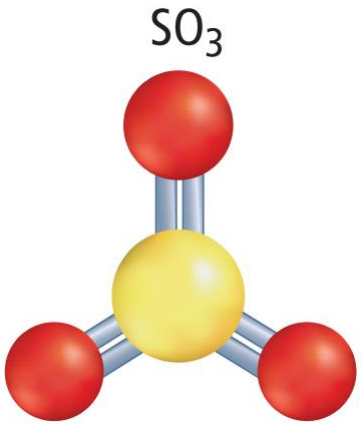
Au, +3; Se, +6; O, −2

b. $Ni(CN)_2$ (nickel(II) cyanide)

Ni, +2; C, +2; N, −3

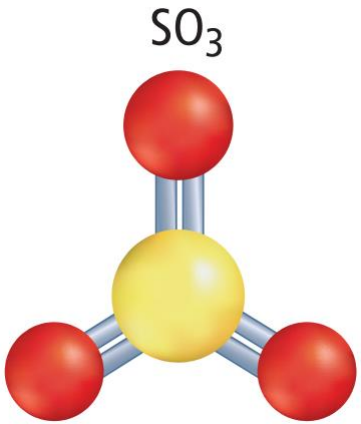
Mastering Problems 50

Explain how the sulfite ion (SO_3^{2-}) differs from sulfur trioxide (SO_3), shown in the figure.



Mastering Problems 50

Explain how the sulfite ion (SO_3^{2-}) differs from sulfur trioxide (SO_3), shown in the figure.



SO_3^{2-} is a polyatomic ion and the oxidation number of sulfur is +4. SO_3 is a compound and the oxidation number of S in this compound is +6.

يُعدّ SO_3^{2-} أيونًا متعدِّد الذرات، وعدد التأكسد للكبريت فيه يساوي 4+، في حين يُعدّ SO_3 مركّبًا وعدد التأكسد للكبريت فيه يساوي 6+.

Page 154

4.1 Oxidation and Reduction

Mixed Review 72

Determine the oxidation number of the boldface element in each of the following.

a. $\mathbf{O}F_2$

b. $\mathbf{U}O_2^{2+}$

c. $\mathbf{Ru}O_4$

d. $\mathbf{Fe_2}O_3$

Page 156

4.1 Oxidation and Reduction

Mixed Review 72

Determine the oxidation number of the boldface element in each of the following.

a. $\mathbf{O}F_2$

O, +2

b. $\mathbf{U}O_2^{2+}$

U, +6

c. $\mathbf{Ru}O_4$

Ru, +8

d. $\mathbf{Fe}_2O_3$

Fe, +3

Page 156

4.1 Oxidation and Reduction

Mixed Review 73

Identify each of the following changes as either oxidation or reduction.

a. $2Cl^- \longrightarrow Cl_2 + 2e^-$

b. $Na \longrightarrow Na^+ + e^-$

c. $Ca^{-2} + 2e^- \longrightarrow 2Ca$

d. $O_2 + 4e^- \longrightarrow 2O^{2-}$

Mixed Review 73

Identify each of the following changes as either oxidation or reduction.

a. $2Cl^- \longrightarrow Cl_2 + 2e^-$

oxidation

b. $Na \longrightarrow Na^+ + e^-$

oxidation

c. $Ca^{-2} + 2e^- \longrightarrow 2Ca$

reduction

d. $O_2 + 4e^- \longrightarrow 2O^{2-}$

reduction

Mixed Review 74

Use the rules for assigning oxidation numbers to complete **Table 7**.

Oxidation Number Assignemtn

Element	Oxidation number
K in KBr	+1
Br in KBr	
Cl in Cl2	
K in KCl	
Cl in KCl	−1
Br in Br2	0

Mixed Review 74

Use the rules for assigning oxidation numbers to complete **Table 7**.

Oxidation Number Assignemtn

Element	Oxidation number
K in KBr	+1
Br in KBr	−1
Cl in Cl2	0
K in KCl	+1
Cl in KCl	−1
Br in Br2	0

Page 156

4.1 Oxidation and Reduction

Mixed Review 75

Identify the reducing agents in these equations.

a. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

b. $Na_2SO_4 + 4C \rightarrow Na_2S + 4CO$

c. $4IrF_5 + Ir \rightarrow 5IrF_4$

Mixed Review 75

Identify the reducing agents in these equations.

a. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

$\overset{0}{O_2} \rightarrow NO^{-2}\ H_2O$ (reduction)

NH_3

b. $Na_2SO_4 + 4C \rightarrow Na_2S + 4CO$

$\overset{+6}{S}\ O_4^{2-} \rightarrow \overset{2-}{S}^{2-}$ (reduction)

C

c. $4IrF_5 + Ir \rightarrow 5IrF_4$

$\overset{+5}{IrF}_5 \rightarrow \overset{+4}{IrF}_4$ (reduction)

Ir

Page 156

4.1 Oxidation and Reduction

Mixed Review 77

What probably accounts for the different forms of chromium shown?

a

b

Mixed Review 77

What probably accounts for the different forms of chromium shown?

a

b

Chromium has different oxidation states in each image.

Mixed Review 79

Rubies are gemstones made up mainly of aluminum oxide. Their red color comes from a small amount of chromium(III) ions replacing some of the aluminum ions. Write the formula for aluminum oxide, and show the reaction in which an aluminum ion is replaced with a chromium ion. Is this a redox reaction?

Mixed Review 79

Rubies are gemstones made up mainly of aluminum oxide. Their red color comes from a small amount of chromium(III) ions replacing some of the aluminum ions. Write the formula for aluminum oxide, and show the reaction in which an aluminum ion is replaced with a chromium ion. Is this a redox reaction?

Al_2O_3; $Al_2O_3 + 2Cr^{3+} \rightarrow Cr_2O_3 + 2Al^{3+}$; it is not a redox reaction—the oxidation number stays the same

Al_2O_3: $Al_2O_3 + 2Cr^{3+} \rightarrow Cr_2O_3 + 2Al^{3+}$

لا، يُعدّ تفاعل أكسدة واختزال؛ لأنه لا يوجد تغيّر في أعداد التأكسد.

Think Critically 82

The following equations show redox reactions that are sometimes used in the laboratory to generate pure nitrogen gas and pure dinitrogen monoxide gas (nitrous oxide, N2O).

$$NH_4NO_2(s) \rightarrow N_2(g) + 2H_2O(l)$$

$$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(l)$$

Write a sentence telling how the electron transfer taking place in these two reactions differs from that taking place in the reaction below.

$$2AgNO_3 + Zn \rightarrow Zn(NO_3)_2 + 2Ag$$

Think Critically 82

The following equations show redox reactions that are sometimes used in the laboratory to generate pure nitrogen gas and pure dinitrogen monoxide gas (nitrous oxide, N2O).

$$NH_4NO_2(s) \rightarrow N_2(g) + 2H_2O(l)$$

$$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(l)$$

Write a sentence telling how the electron transfer taking place in these two reactions differs from that taking place in the reaction below.

$$2AgNO_3 + Zn \rightarrow Zn(NO_3)_2 + 2Ag$$

In the first two reactions, nitrogen is oxidized and reduced. The third reaction involves redox between two different elements.

في التفاعلين الأولين تأكسَد النيتروجين واختُزِل، أمّا في التفاعل الثالث فقد حدث تفاعل الأكسدة والاختزال بين عنصرين مختلفين.

Think Critically 84

Consider the fact that all of the following are stable compounds. What can you infer about the oxidation state of phosphorus in its compounds?

$$PH_3,\ PCl_3,\ P_2H_4,\ PCl_5,\ H_3PO_4,\ Na_3PO_3$$

Think Critically 84

Consider the fact that all of the following are stable compounds. What can you infer about the oxidation state of phosphorus in its compounds?

$$PH_3,\ PCl_3,\ P_2H_4,\ PCl_5,\ H_3PO_4,\ Na_3PO_3$$

Phosphorus has several oxidation states (−3, −2, +3, +5) that make phosphorus very flexible when combining with nonmetals.

للفوسفـور حالات تأكسد متعـدِّدة (3−، 2−، 3+، 5+) ممّا يجعله مرنًا عند اتحاده باللافلزات.

Multiple Choice 1

Which is NOT a reducing agent in a redox reaction?

a) the substance oxidized

b) the electron acceptor

c) the less-electronegative substance

d) the electron donor

Multiple Choice 1

Which is NOT a reducing agent in a redox reaction?

a) the substance oxidized

b) the electron acceptor

c) the less-electronegative substance

d) the electron donor

Page 158

4.1 Oxidation and Reduction

Multiple Choice 3

The reaction between sodium iodide and chlorine is shown below.

$$2NaI(aq) + Cl_2(aq) \rightarrow 2NaCl(aq) + I_2(aq)$$

The oxidation state of sodium remains unchanged for which reason?

a) Na^+ is a spectator ion.

b) Na^+ cannot be reduced.

c) Na is an uncombined element.

d) Na^+ is a monatomic ion.

Multiple Choice 3

The reaction between sodium iodide and chlorine is shown below.

$$2NaI(aq) + Cl_2(aq) \rightarrow 2NaCl(aq) + I_2(aq)$$

The oxidation state of sodium remains unchanged for which reason?

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c) Na is an uncombined element.

d) Na^+ is a monatomic ion.

Multiple Choice 4

The reaction between sodium iodide and chlorine is shown below.

$$2NaI(aq) + Cl_2(aq) \rightarrow 2NaCl(aq) + I_2(aq)$$

Which of the following is the oxidizing agent in the reaction?

a) Cl_2

b) I_2

c) NaCl

d) NaI

Page 158

4.1 Oxidation and Reduction

Multiple Choice 4

The reaction between sodium iodide and chlorine is shown below.

$$2NaI(aq) + Cl_2(aq) \rightarrow 2NaCl(aq) + I_2(aq)$$

Which of the following is the oxidizing agent in the reaction?

a) Cl_2

b) I_2

c) NaCl

d) NaI

Page 158

4.1 Oxidation and Reduction

Balancing Redox Equations

Chapter 4

Lesson 2 – Revision Paper

Section Summary

- Redox equations in which the same element appears in multiple reactants and products can be difficult to balance using the conventional method.

- To balance equations for reactions in an acidic solution, add hydrogen ions and water molecules.

- To balance equations for reactions in a basic solution, add hydroxide ions and water molecules.

- A half-reaction is one of the two parts of a redox reaction.

Example 4

Balance the following redox equation.

$$ClO_4^-(aq) + Br^-(aq) \rightarrow Cl^-(aq) + Br_2(g) \text{ (in acidic solution)}$$

Example 4

Balance the following redox equation.

$$ClO_4^-(aq) + Br^-(aq) \rightarrow Cl^-(aq) + Br_2(g) \text{ (in acidic solution)}$$

Assign oxidation numbers to all elements in the equation.

+7 −2 −1 −1 0

$ClO_4^-(aq) + Br^-(aq) \rightarrow Cl^-(aq) + Br_2(g)$ (in acidic solution)

Use the rules in Table 2.

Identify which atoms or ions are oxidized and which are reduced.

Br is oxidized.
Cl is reduced.

The oxidation number of bromine increases from −1 to 0. The oxidation number of chlorine decreases from +7 to −1.

Determine the changes in oxidation number for the atoms or ions that are oxidized and reduced.

Changes in oxidation number:

Br +1
Cl −8

Bromine loses electrons. It is oxidized.
Chlorine gains electrons. It is reduced.

Make the changes in oxidation number equal in magnitude by adjusting coefficients in the equation.

$ClO_4^-(aq) + 8Br^-(aq) \rightarrow Cl^-(aq) + 4Br_2(g)$ (in acidic solution)

Because the change in oxidation number of Br is +1, you must add the coefficient 8 to balance the equation. $4Br_2$ represents 8 Br atoms to balance the $8Br^-$ on the left side.

Add enough hydrogen ions and water molecules to the equation to balance the oxygen atoms on both sides.

$ClO_4^-(aq) + 8Br^-(aq) + 8H^+(aq) \rightarrow Cl^-(aq) + 4Br_2(g) + 4H_2O(l)$

Because you know the reaction takes place in acid solution, you can add H^+ ions on the left side of the equation.

Applications 19

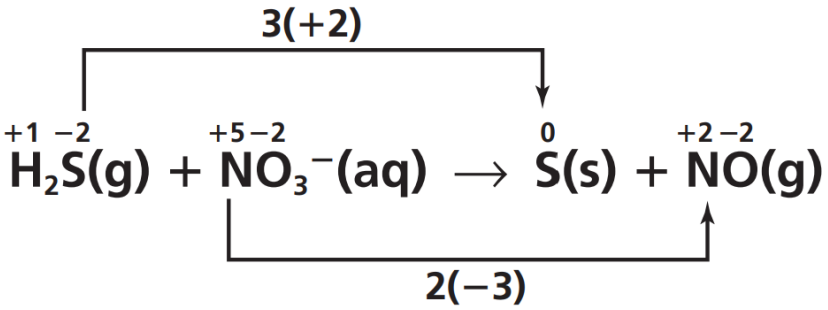
Balance these redox equations.

19. $H_2S(g) + NO_3^-(aq) \longrightarrow S(s) + NO(g)$(in acid solution)

Applications 19

Balance these redox equations.

19. $H_2S(g) + NO_3^-(aq) \longrightarrow S(s) + NO(g)$(in acid solution)



$2H^+(aq) + 3H_2S(g) + 2NO_3^-(aq) \longrightarrow 3S(s) + 2NO(g) + 4H_2O(l)$

oxidation number of S increases from −2 to 0;
oxidation number of N decreases from +5 to +2

Applications 20

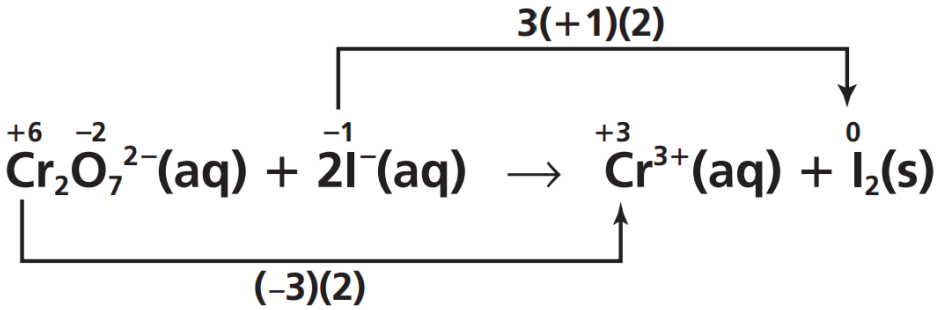
Balance these redox equations.

20. $Cr_2O_7^{2-}(aq) + I^-(aq) \longrightarrow Cr^{3+}(aq) + I_2(s)$(in acid solution)

Applications 20

Balance these redox equations.

20. $Cr_2O_7^{2-}(aq) + I^-(aq) \longrightarrow Cr^{3+}(aq) + I_2(s)$(in acid solution)



$14H^+(aq) + Cr_2O_7^{2-}(aq) + 6I^- \longrightarrow 2Cr^{3+}(aq) + 3I_2(s) + 7H_2O(l)$

oxidation number of I increases from −1 to 0;
oxidation number of Cr decreases from +6 to +3

Applications 21

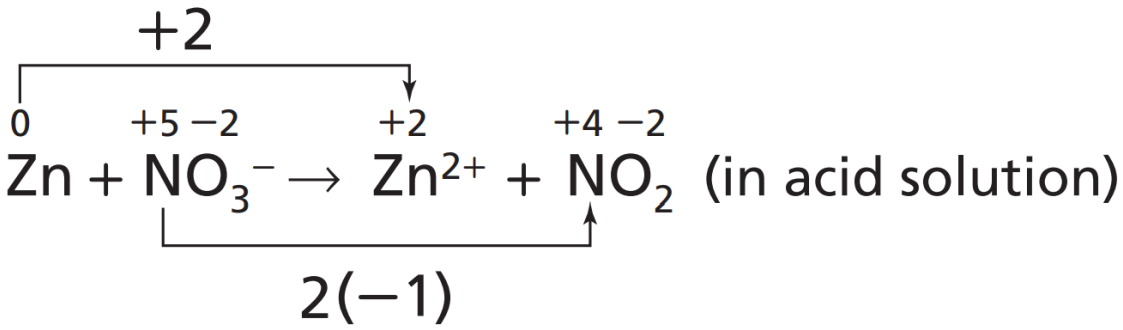
Balance these redox equations.

21. $Zn + NO_3^- \longrightarrow Zn^{2+} + NO_2$(in acid solution)

Applications 21

Balance these redox equations.

21. $Zn + NO_3^- \longrightarrow Zn^{2+} + NO_2$(in acid solution)



oxidation number of Zn increases from 0 to +2;
oxidation number of N decreases from +5 to +4

$$Zn + 2NO_3^- + 4H^+ \rightleftarrows Zn^{2+} + 2NO_2 + 2H_2O$$

Applications 22

Balance these redox equations.

22. Challenge $I^-(aq) + MnO_4^-(aq) \longrightarrow I_2(s) + MnO_2(s)$ (in basic solution)

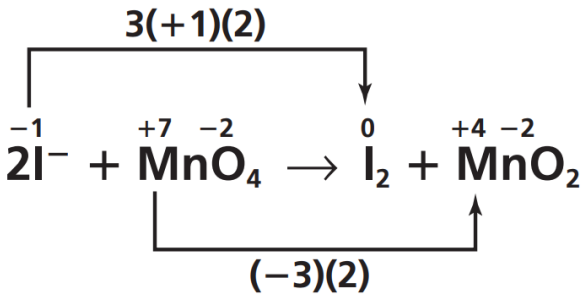
Applications 22

Balance these redox equations.

22. Challenge $I^-(aq) + MnO_4^-(aq) \longrightarrow I_2(s) + MnO_2(s)$ (in basic solution)

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

oxidation number of I increases from −1 to 0;
oxidation number of Mn decreases from +7 to +4



Example 5

Balance the redox equation for the reaction below using half-reactions.

$$KMnO_4(aq) + SnCl_2(aq) + HCl(aq) \rightarrow MnCl_2(aq) + SnCl_4(aq) + H_2O(l) + KCl(aq)$$

Example 5

Balance the redox equation for the reaction below using half-reactions.

$$KMnO_4(aq) + SnCl_2(aq) + HCl(aq) \rightarrow MnCl_2(aq) + SnCl_4(aq) + H_2O(l) + KCl(aq)$$

Write the unbalanced, net ionic equation for the reaction.

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

Eliminate coefficients, spectator ions, and state symbols.

Write incomplete equations for the oxidation and reduction half-reactions, including oxidation numbers.

$\overset{+2}{Sn^{2+}} \rightarrow \overset{+4}{Sn^{4+}}$ (oxidation)

$\overset{+7}{MnO_4^-} \rightarrow \overset{+2}{Mn^{2+}}$ (reduction)

Use the rules in Table 2 and Table 6.

Balance the atoms and charges in the half-reactions.

$Sn^{2+} \rightarrow Sn^{4+} + 2e^-$ (oxidation)

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

In an acid solution, H_2O molecules are available in abundance and can be used to balance oxygen atoms in the half-reactions; H^+ ions are readily available and can be used to balance the charge.

Adjust the coefficients so that the number of electrons lost in oxidation (2) equals the number of electrons gained in reduction (5).

$5Sn^{2+} \rightarrow 5Sn^{4+} + 10e^-$ (oxidation)

$10e^- + 16H^+ + 2MnO_4^- \rightarrow 2Mn^{2+} + 8H_2O$ (reduction)

The least common multiple of 2 and 5 is 10. Cross-multiplying gives the balanced oxidation and reduction half-reactions.

Add the balanced half-reactions and simplify by canceling or reducing like terms on both sides of the equation.

$5Sn^{2+} + 10e^- + 16H^+ + 2MnO_4^- \rightarrow 5Sn^{4+} + 10e^- + 2Mn^{2+} + 8H_2O$

$5Sn^{2+} + 16H^+ + 2MnO_4^- \rightarrow 5Sn^{4+} + 2Mn^{2+} + 8H_2O$

Restore state descriptions and return spectator ions (K^+ and Cl^-).

$5SnCl_2(aq) + 16HCl(aq) + 2KMnO_4(aq) \rightarrow 5SnCl_4(aq) + 2MnCl_2(aq) + 8H_2O(l) + 2KCl(aq)$

Add K^+ ions to the two MnO_4^- ions on the left and add two K^+ ions on the right. Add Cl^- ions to the Sn^{2+} and H^+ ions on the left and to the Sn^{4+}, Mn^{2+}, and K^+ ions on the right.

Applications 23

Use the half-reaction method to balance the redox equations. Begin by writing the oxidation and reduction half-reactions. Leave the balanced equation in ionic form.

23. $Cr_2O_7^{2-}(aq) + I^-(aq) \longrightarrow Cr^{3+}(aq) + I_2(s)$(in acid solution)

Applications 23

Use the half-reaction method to balance the redox equations. Begin by writing the oxidation and reduction half-reactions. Leave the balanced equation in ionic form.

23. $Cr_2O_7^{2-}(aq) + I^-(aq) \longrightarrow Cr^{3+}(aq) + I_2(s)$(in acid solution)

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

$14H^+(aq) + 6e^- + Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$ (reduction)

Multiply oxidation half-reaction by 3 and add to reduction half-reaction

$14H^+(aq) + 6e^- + Cr_2O_7^{2-}(aq) + 6I^-(aq) \rightarrow 3I_2(s) + 2Cr^{3-}(aq) + 7H_2O(l) + 6e^-$

$14H^+(aq) + Cr_2O_7^{2-}(aq) + 6I^-(aq) \rightarrow 3I_2(s) + 2Cr^{3-}(aq) + 7H_2O(l)$

Applications 24

Use the half-reaction method to balance the redox equations. Begin by writing the oxidation and reduction half-reactions. Leave the balanced equation in ionic form.

$$Mn^{2+}(aq) + BiO_3^-(aq) \rightarrow MnO_4^-(aq) + Bi^{2+}(aq) \text{ (in acidic solution)}$$

Applications 24

Use the half-reaction method to balance the redox equations. Begin by writing the oxidation and reduction half-reactions. Leave the balanced equation in ionic form.

$$Mn^{2+}(aq) + BiO_3^-(aq) \rightarrow MnO_4^-(aq) + Bi^{2+}(aq) \text{ (in acidic solution)}$$

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

$BiO_3^-(aq) + 3e^- + 6H^+(aq) \rightarrow Bi^{2+}(aq) + 3H_2O(l)$ (reduction)

Multiply oxidation half-reaction by 3. Multiply reduction half-reaction by 5 and add to oxidation half-reaction.

$3Mn^{2+}(aq) + 12H_2O(l) + 5BiO_3^-(aq) + 15e^- + 30H^+(aq) \rightarrow 3MnO_4^-(aq) + 15e^- + 24H^+(aq) + 5Bi^{2+}(aq) + 15H_2O(l)$

$3Mn^{2+}(aq) + 5BiO_3^-(aq) + 6H^+(aq) \rightarrow 3MnO_4^-(aq) + 5Bi^{2+}(aq) + 3H_2O(l)$

Page 149

Applications 25

CHALLENGE Use the half-reaction method to balance the redox equations. Begin by writing the oxidation and reduction half-reactions. Leave the balanced equation in ionic form.

$$N_2O(g) + ClO^-(aq) \rightarrow NO_2^-(aq) + Cl^-(aq) \text{ (in basic solution)}$$

Applications 25

CHALLENGE Use the half-reaction method to balance the redox equations. Begin by writing the oxidation and reduction half-reactions. Leave the balanced equation in ionic form.

$$N_2O(g) + ClO^-(aq) \longrightarrow NO_2^-(aq) + Cl^-(aq) \text{ (in basic solution)}$$

$6OH^-(aq) + N_2O(g) \longrightarrow 2NO_2^-(aq) + 4e^- + 3H_2O(l)$(oxidation)

$ClO^-(aq) + 2e^- + H_2O(l) \longrightarrow Cl^-(aq) + 2OH^-$(aq)(reduction)

reduction half-reaction by 2 and add to oxidation half-reaction.

$6OH^-(aq) + N_2O(g) + 2ClO^-(aq) + 4e^- + 2H_2O(l) \longrightarrow 2NO_2^-(aq) + 4e^- + 3H_2O(l) + 2Cl^-(aq) + 4OH^-(aq)$

$N_2O(g) + 2ClO^-(aq) + 2OH^-(aq) \longrightarrow 2NO_2^-(aq) + 2Cl^-(aq) + H_2O(l)$

Page 149

4.2 Balancing Redox Equations

Apply the Strategy

Balance the following equation.

$$P_4(s) + H_2O(l) \rightarrow PH_3(g) + H_2PO_2(aq)$$

Apply the Strategy

Balance the following equation.

$$P_4(s) + H_2O(l) \rightarrow PH_3(g) + H_2PO_2(aq)$$

$$P_4 + H_2O \rightarrow PH_3 + H_2PO_2$$
$$5P_4 + 24H_2O \rightarrow 8PH_3 + 12H_2PO_2$$

Review 26

Explain how changes in oxidation number are related to the electrons transferred in a redox reaction. How are the changes related to the processes of oxidation and reduction.

Review 26

Explain how changes in oxidation number are related to the electrons transferred in a redox reaction. How are the changes related to the processes of oxidation and reduction.

Because the nucleus (specifically, number of protons) never changes during this type of reaction, whenever there is a transfer of electrons from one atom to another atom, there is a change in the net charge of that species. Oxidation increases the oxidation number; reduction reduces it.

عندما يحدث انتقال للإلكترونات من ذرة إلى أخرى خلال تفاعلات الأكسدة والاختزال يحدث تغير في الشحنة الكلية لهذه الذرات؛ وذلك لأن النواة، وبخاصة عدد البروتونات فيها، لا تتغيّر خلال هذا النوع من التفاعلات أبدًا.

Review 27

Describe why it is important to know the conditions under which an aqueous oxidation-reduction reaction takes place in order to balance the ionic equation for the reaction.

Review 27

Describe why it is important to know the conditions under which an aqueous oxidation-reduction reaction takes place in order to balance the ionic equation for the reaction.

It is important to know whether H^+ or OH^- ions are available to balance the equation.

من المهم معرفة وجود أيونات الهيدروجين وأيونات الهيدروكسيد لوزن المعادلة.

Review 29

An oxidation half-reaction shows the _____ of electrons from an atom that undergoes a/an _____ in oxidation number.

a) production – increase

b) production – decrease

c) combining – increase

d) combining – decrease

Review 29

An oxidation half-reaction shows the _____ of electrons from an atom that undergoes a/an _____ in oxidation number.

a) production – increase

b) production – decrease

c) combining – increase

d) combining – decrease

Review 29

A reduction half-reaction shows the _____ of electrons with an atom that undergoes a _____ in oxidation number.

a) combining – increase

b) combining – decrease

c) production – increase

d) production – decrease

Review 29

A reduction half-reaction shows the _____ of electrons with an atom that undergoes a _____ in oxidation number.

a) combining – increase

b) combining – decrease

c) production – increase

d) production – decrease

Review 30

Write the oxidation and reduction half-reactions for the redox equation.

$$Pb(s) + Pd(NO_3)_2(aq) \longrightarrow Pb(NO_3)_2(aq) + Pd(s)$$

Page 150

4.2 Balancing Redox Equations

Review 30

Write the oxidation and reduction half-reactions for the redox equation.

$$Pb(s) + Pd(NO_3)_2(aq) \longrightarrow Pb(NO_3)_2(aq) + Pd(s)$$

oxidation: $Pb \longrightarrow Pb^{2+} + 2e^-$

reduction: $Pd^{2+} + 2e^- \longrightarrow Pd$

Review 31

The oxidation half-reaction of a redox reaction is $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$, and the reduction half-reaction is $Au^{3+} + 3e^- \rightarrow Au$. What minimum numbers of tin(II) ions and gold(III) ions would have to react in order to have zero electrons left over?

Review 31

The oxidation half-reaction of a redox reaction is $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$, and the reduction half-reaction is $Au^{3+} + 3e^- \rightarrow Au$. What minimum numbers of tin(II) ions and gold(III) ions would have to react in order to have zero electrons left over?

$$\mathbf{Sn^{2+} \rightarrow Sn^{4+} + 2e^-}$$

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

$$\mathbf{3Sn^{2+} + 2Au^{3+} + \cancel{6e^-} \rightarrow 3Sn^{4+} + 2Au + \cancel{6e^-}}$$

Three Sn^{2+} ions; two Au^{3+} ions

Review 32

Balance the following equations.

a. $HClO_3(aq) \longrightarrow ClO_2(g) + HClO_4(aq) + H_2O(l)$

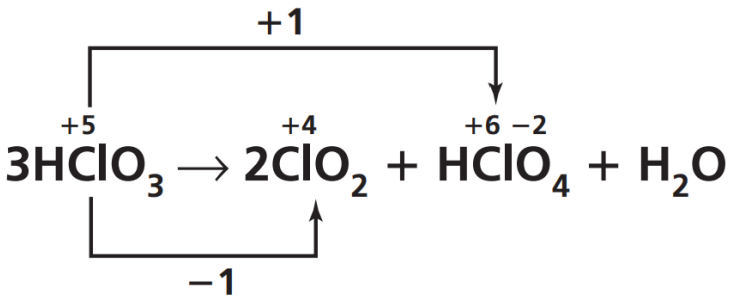
Page 150

4.2 Balancing Redox Equations

Review 32

Balance the following equations.

a. $HClO_3(aq) \longrightarrow ClO_2(g) + HClO_4(aq) + H_2O(l)$



$3HClO_3 \longrightarrow 2ClO_2 + HClO_4 + H_2O$

Page 150

4.2 Balancing Redox Equations

Review 32

Balance the following equations.

b. $H_2SeO_3(aq) + HClO_3(aq) \longrightarrow H_2SeO_4(aq) + Cl_2(g) + H_2O(l)$

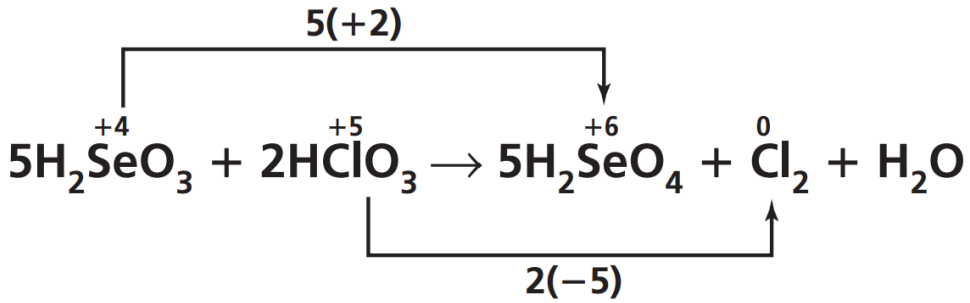
Page 150

4.2 Balancing Redox Equations

Review 32

Balance the following equations.

b. $H_2SeO_3(aq) + HClO_3(aq) \longrightarrow H_2SeO_4(aq) + Cl_2(g) + H_2O(l)$



$$5H_2SeO_3 + 2HClO_3 \longrightarrow 5H_2SeO_4 + Cl_2 + H_2O$$

Page 150

4.2 Balancing Redox Equations

Review 32

Balance the following equations.

c. $Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \longrightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$ (in acid solution)

Review 32

Balance the following equations.

c. $Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \longrightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$ (in acid solution)

$Cr2O_7^{2-} + 6Fe^{2+} + 14H^{+} \longrightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$

Mastering Concepts 51

Compare and contrast balancing redox equations in acidic and basic solutions.

Mastering Concepts 51

Compare and contrast balancing redox equations in acidic and basic solutions.

In a redox reaction that takes place in an acidic solution, H^+ and H_2O can participate in the reaction as either reactants or products. In a basic solution, a redox reaction may involve OH^- and H_2O as either reactants or products.

يمكن لـ H^+ و H_2O أن تشارك في تفاعلات الأكسدة والاختزال التي تحدث في المحاليل الحمضية، إمّا بوصفها متفاعلات أو نواتج. ويتضمَّن تفاعل الأكسدة والاختزال في المحلول القاعدي OH^- و H_2O إمّا على صورة متفاعلات أو نواتج.

Mastering Concepts 52

Explain why writing hydrogen ions as $\mathbf{H^+}$ in redox reactions represents a simplification and not how they exist.

Mastering Concepts 52

Explain why writing hydrogen ions as H^+ in redox reactions represents a simplification and not how they exist.

In aqueous solution, hydrogen ions combine with water in their hydrated form, the hydronium ions (H_3O^+) and are never present as H^+. However, they are sometimes shown as H^+ to simplify the chemical equation that is written.

تتّحد أيونات الهيدروجين بالماء في المحاليل المائية في شكلها المائي، أيونات الهيدرونيوم H_3O^+، ولا يمكن أن توجد في صورة H^+. ولكنها تُكتَب في بعض الأحيان في صورة H^+ لتبسيط المعادلة الكيميائية المكتوبة.

Page 154

4.2 Balancing Redox Equations

Mastering Concepts 53

Before you attempt to balance the equation for a redox reaction, why do you need to know whether the reaction takes place in acidic or basic solution?

Mastering Concepts 53

Before you attempt to balance the equation for a redox reaction, why do you need to know whether the reaction takes place in acidic or basic solution?

The type of solution determines whether H^+ or OH^- ions are available to balance the redox equation.

يحدد نوع المحلول سواء كان H^+ أو أيونات OH^- متاحة لوزن معادلة الاختزال.

Mastering Concepts 54

_____ is one that is present in the same stoichiometry on both sides of a redox reaction equation. They are not changed during a reaction, so they can be eliminated from the equation.

a) Species

b) Spectator ion

c) Half ion

d) Oxidation ion

Mastering Concepts 54

_____ is one that is present in the same stoichiometry on both sides of a redox reaction equation. They are not changed during a reaction, so they can be eliminated from the equation.

a) Species

b) Spectator ion

c) Half ion

d) Oxidation ion

الأيونات المتفرّجة هي الأيونات التي توجد في الحسابات الكيميائية على طرفي معادلة الأكسدة والاختزال بالمقدار نفسه. لكنها لا تتغيّر في أثناء التفاعل، لذا يمكن حذفها من المعادلة.

Page 154

4.2 Balancing Redox Equations

Mastering Concepts 55

A/An _____ is any kind of chemical unit involved in the redox process. It can be an ion, molecule, or a free atom.

a) Reduction ion

b) Species

c) Half-reaction

d) Oxidation ion

Mastering Concepts 55

A/An _____ is any kind of chemical unit involved in the redox process. It can be an ion, molecule, or a free atom.

a) Reduction ion

b) Species

c) Half-reaction

d) Oxidation ion

المادة أيّ صنف من الوحدات الكيميائية توجد في عمليات الأكسدة أو الاختزال، وقد تكون أيونًا أو جزيئًا، أو ذرات حرة.

Page 154

4.2 Balancing Redox Equations

Mastering Concepts 56

Is the following equation balanced? Explain.

$$Fe(s) + Ag^{+}(aq) \rightarrow Fe^{2+}(aq) + Ag(s)$$

Page 154

4.2 Balancing Redox Equations

Mastering Concepts 56

Is the following equation balanced? Explain.

$$Fe(s) + Ag^{+}(aq) \rightarrow Fe^{2+}(aq) + Ag(s)$$

The total charge on the left-hand side does not equal the total charge on the right-hand side.

لا تساوي الشحنةُ الكلية في الجهة اليسرى الشحنةَ الكلية في الجهة اليمنى.

Mastering Concepts 57

Does the following equation represent a reduction or an oxidation process? Explain your answer.

$$Zn^{2+} + 2e^- \rightarrow Zn$$

Mastering Concepts 57

Does the following equation represent a reduction or an oxidation process? Explain your answer.

$$Zn^{2+} + 2e^- \rightarrow Zn$$

reduction; Electrons are gained and the oxidation number for Zn decreases.

عملية اختزال؛ إذ تُكتَسَب الإلكترونات، ويقلّ عدد تأكسد الخارصين Zn.

Page 154

4.2 Balancing Redox Equations

Mastering Concepts 58

Describe what is happening to electrons in each half reaction of a redox process.

Page 155

4.2 Balancing Redox Equations

Mastering Concepts 58

Describe what is happening to electrons in each half reaction of a redox process.

Electrons are accepted by a species during the reduction half-reaction, and electrons are lost from species during an oxidation half-reaction.

تُكتَسَب الإلكترونات من قبل بعض المواد خلال نصف تفاعل الاختزال، وتُفقَد الإلكترونات من بعض المواد خلال نصف تفاعل الأكسدة.

Mastering Problems 6o

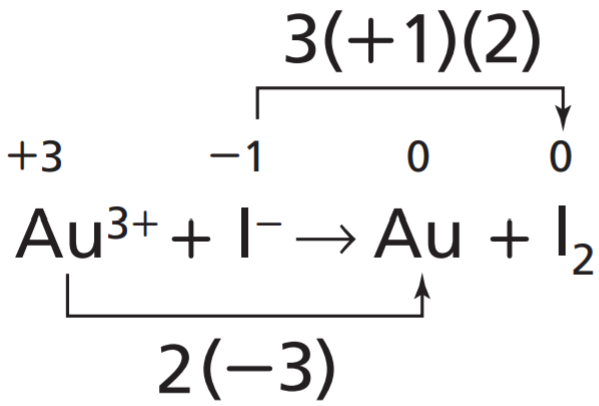
Balance these net ionic equations for redox reactions.

a. $Au^{3+}(aq) + I^{-}(aq) \longrightarrow Au(s) + I_2(s)$

Mastering Problems 6o

Balance these net ionic equations for redox reactions.

a. $Au^{3+}(aq) + I^{-}(aq) \longrightarrow Au(s) + I_2(s)$



$2Au^{3+}(aq) + 6I^{-}(aq) \longrightarrow 2Au(s) + 3I_2(s)$

Page 155

4.2 Balancing Redox Equations

Mastering Problems 6o

Balance these net ionic equations for redox reactions.

b. $Ce^{4+}(aq) + Sn^{2+}(aq) \longrightarrow Ce^{3+}(aq) + Sn^{4+}(aq)$

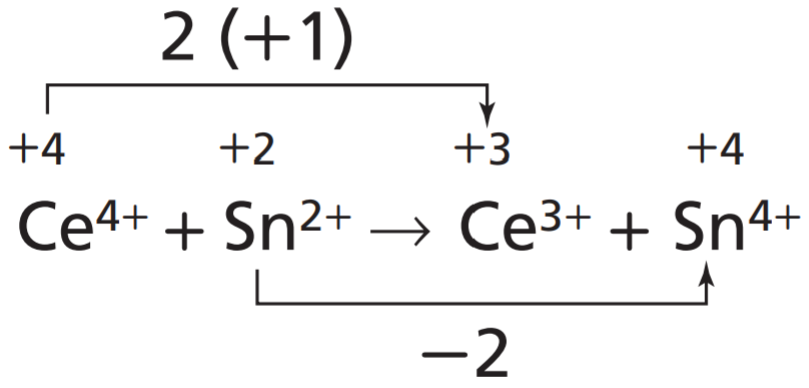
Page 155

4.2 Balancing Redox Equations

Mastering Problems 6o

Balance these net ionic equations for redox reactions.

b. $Ce^{4+}(aq) + Sn^{2+}(aq) \longrightarrow Ce^{3+}(aq) + Sn^{4+}(aq)$



$$2Ce^{4+} + Sn^{2+} \longrightarrow Sn^{4+} + 2Ce^{3+}$$

Page 155

4.2 Balancing Redox Equations

Mastering Problems 61

Balance the following ionic redox equations.

a. $\mathrm{Al} + \mathrm{I_2} \longrightarrow \mathrm{Al^{3+}} + \mathrm{I^-}$

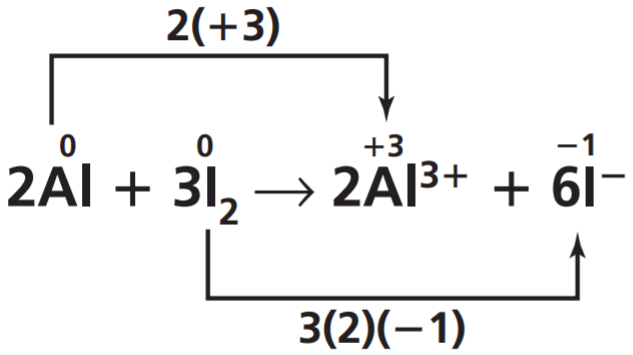
Page 155

4.2 Balancing Redox Equations

Mastering Problems 61

Balance the following ionic redox equations.

a. $Al + I_2 \longrightarrow Al^{3+} + I^-$



$$2Al + 3I_2 \longrightarrow 2Al^{3+} + 6I^-$$

Page 155

4.2 Balancing Redox Equations

Mastering Problems 61

Balance the following ionic redox equations.

b. $MnO_2 + Br^- \longrightarrow Mn^{2+} + Br_2$ (in acid solution)

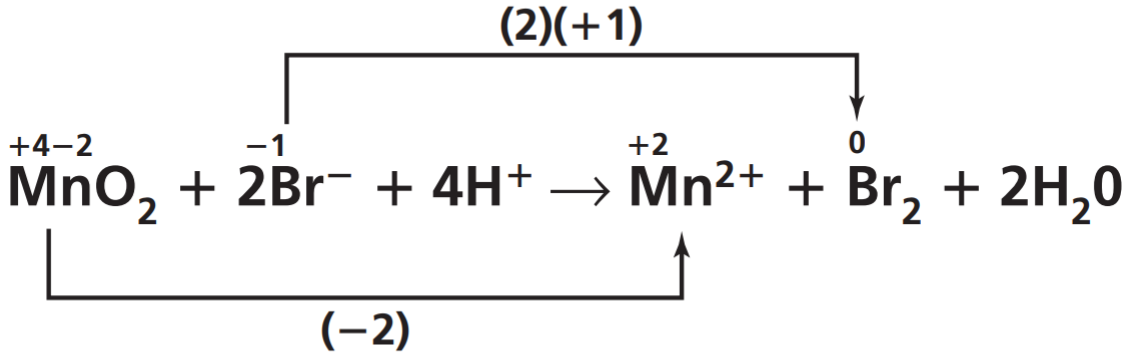
Page 155

4.2 Balancing Redox Equations

Mastering Problems 61

Balance the following ionic redox equations.

b. $MnO_2 + Br^- \longrightarrow Mn^{2+} + Br_2$ (in acid solution)



$$MnO_2 + 2Br^- + 4H^+ \longrightarrow Mn^{2+} + Br_2 + 2H_2O$$

Mastering Problems 63

The mineral corundum is comprised of aluminum oxide (Al_2O_3) and is dull gray. Sapphire is mostly aluminum oxide, but it contains small amounts of Fe (2+) and Ti (4+). The color of sapphire results from an electron transfer from Fe (2+) to Ti (4+). Write an equation that describes the reaction that occurs resulting in the mineral on the right. What are the oxidizing and reducing agents?

Mastering Problems 63

The mineral corundum is comprised of aluminum oxide (Al_2O_3) and is dull gray. Sapphire is mostly aluminum oxide, but it contains small amounts of Fe (2+) and Ti (4+). The color of sapphire results from an electron transfer from Fe (2+) to Ti (4+). Write an equation that describes the reaction that occurs resulting in the mineral on the right. What are the oxidizing and reducing agents?

$Fe^{2+} + Ti^{4+} \rightarrow Fe^{3+} + Ti^{3+}$; Fe is the reducing agent, Ti is the oxidizing agent

$$Fe^{2+} + Ti^{4+} \rightarrow Fe^{3+} + Ti^{3+}$$

يُعدّ الحديد Fe العامل المُختَزِل، في حين يُعدّ التيتانيوم Ti العامل المؤكسِد.

Mastering Problems 64

Write the oxidation and reduction half-reactions represented in each of these redox equations. Write the half-reactions in net ionic form if they occur in aqueous solution.

a. $PbO(s) + NH_3(g) \longrightarrow N_2(g) + H_2O(l) + Pb(s)$

b. $I_2(s) + Na_2S_2O_3(aq) \longrightarrow Na_2S_2O_4(aq) + NaI(aq)$

c. $Sn(s) + 2HCl(aq) \longrightarrow SnCl_2(aq) + H_2(g)$

Page 155

4.2 Balancing Redox Equations

Mastering Problems 64

Write the oxidation and reduction half-reactions represented in each of these redox equations. Write the half-reactions in net ionic form if they occur in aqueous solution.

a. $PbO(s) + NH_3(g) \longrightarrow N_2(g) + H_2O(l) + Pb(s)$

$NH_3(g) \longrightarrow N_2(g) + 3e^-$ oxidation

$PbO(s) + 2e^- \longrightarrow Pb(s)$ reduction

b. $I_2(s) + Na_2S_2O_3(aq) \longrightarrow Na_2S_2O_4(aq) + NaI(aq)$

$I_2(s) + 2e^- \longrightarrow 2I^-(aq)$ reduction

$2S_2O_3^{2-}(aq) \longrightarrow 2S_2O_4^{2-}(aq) + 2e^-$ oxidation

c. $Sn(s) + 2HCl(aq) \longrightarrow SnCl_2(aq) + H_2(g)$

$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^-$ oxidation

$2H^+(aq) + 2e^- \longrightarrow H_2(g)$ reduction

Page 155

4.2 Balancing Redox Equations

Mastering Problems 65

Write the two half-reactions that make up the following balanced redox reaction.

$$3H_2C_2O_4 + 2HAsO_2 \longrightarrow 6CO_2 + 2As + 4H_2O$$

Page 155

4.2 Balancing Redox Equations

Mastering Problems 65

Write the two half-reactions that make up the following balanced redox reaction.

$$3H_2C_2O_4 + 2HAsO_2 \longrightarrow 6CO_2 + 2As + 4H_2O$$

$$H_2C_2O_4 \longrightarrow 2CO_2 + 2H^+ + 2e^-$$

$$HAsO_2 + 3H^+ + 3e^- \longrightarrow As + 2H_2O$$

Page 155

4.2 Balancing Redox Equations

Mastering Problems 66

Label each half-reaction as reduction or oxidation.

a. $Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + e^-$

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

c. $2H^+ + 2e^- \longrightarrow H_2$

d. $F_2 \longrightarrow 2F^- + 2e$

Page 155

4.2 Balancing Redox Equations

Mastering Problems 66

Label each half-reaction as reduction or oxidation.

a. $Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + e^-$

oxidation

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

reduction

c. $2H^+ + 2e^- \longrightarrow H_2$

reduction

d. $F_2 \longrightarrow 2F^- + 2e$

oxidation

Mastering Problems 67

When **solid copper** is put into a solution of **silver nitrate**, as shown in Figure 12, **silver metal** appears and **blue copper(II) nitrate** forms. Write the corresponding, unbalanced chemical equation. Next, determine the oxidation state of each element in the equation. Write the two half-reactions, labeling each as oxidation or reduction. Finally, write a balanced equation for the reaction.

Figure 12

Mastering Problems 67

When **solid copper** is put into a solution of **silver nitrate**, as shown in Figure 12, **silver metal** appears and **blue copper(II) nitrate** forms. Write the corresponding, unbalanced chemical equation. Next, determine the oxidation state of each element in the equation. Write the two half-reactions, labeling each as oxidation or reduction. Finally, write a balanced equation for the reaction.

Unbalanced: $AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$

Reactant oxidation states: Ag, +1; N, +5; O, −2; Cu, 0

Product oxidation states: Ag, 0; N, +5; O, −2; Cu, +2

Oxidation half-reaction: $Cu \rightarrow Cu^{2+} + 2e^-$

Reduction half-reaction: $e^- + Ag^{+1} \rightarrow Ag$

$2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

■ **Figure 12**

Page 155

4.2 Balancing Redox Equations

Mastering Problems 69

Use the half-reaction method to balance these equations for redox reactions. Add water molecules and hydrogen ions or hydroxide ions as needed.

a. $NH_3(g) + NO_2(g) \longrightarrow N_2(g) + H_2O(l)$

Mastering Problems 69

Use the half-reaction method to balance these equations for redox reactions. Add water molecules and hydrogen ions or hydroxide ions as needed.

a. $NH_3(g) + NO_2(g) \longrightarrow N_2(g) + H_2O(l)$

$2\,\overset{-3}{N}H_3 + N_2 \longrightarrow 6e^-$

$2\,\overset{+4}{N}O_2 + 8e^- \overset{0}{N}_2$

$4(2NH_3 \longrightarrow N_2 + 6H^+ + 6e^-)$

$3(2NO_2 + 8H^+ + 8e^- \longrightarrow N_2 + 4H_2O)$

$8NH_3 + 6NO_2 + \cancel{24H^+} + \cancel{24e^-} \longrightarrow 4N_2 + \cancel{24H^+} + 3N_2 + 12H_2O + \cancel{24e^-}$

$8NH_3 + 6NO_2 \longrightarrow 7N_2 + 12H_2O$

Page 155

4.2 Balancing Redox Equations

Mastering Problems 69

Use the half-reaction method to balance these equations for redox reactions. Add water molecules and hydrogen ions or hydroxide ions as needed.

b. $Br_2 \longrightarrow Br^- + BrO_3^-$ (in basic solution)

Mastering Problems 69

Use the half-reaction method to balance these equations for redox reactions. Add water molecules and hydrogen ions or hydroxide ions as needed.

b. $Br_2 \longrightarrow Br^- + BrO_3^-$ (in basic solution)

$\overset{0}{Br_2} \longrightarrow 2\overset{+5}{Br}\,O_3^- + 10e^-$

$\overset{0}{Br_2} + 2e^- \longrightarrow 2Br^-$

$Br_2 + 12OH^- \longrightarrow 2BrO_3^- + 6H_2O + 10e^-$

$5\,(Br_2 + 10e^- \longrightarrow 10Br^-)$

$\frac{1}{2}(6Br_2 + 12OH^- + 10e^- \longrightarrow 2BrO_3^- + 6H_2O + 10Br^- + 10e^-)$

$3Br_2 + 6OH^- \longrightarrow 5Br^- + BrO_3^- + 3H_2O$

Page 155

4.2 Balancing Redox Equations

Mastering Problems 71

Use the half-reaction method to balance these equations. Add water molecules and hydrogen ions or hydroxide ions as needed.

a. $PbO(s) + NH_3(g) \longrightarrow N_2(g) + H_2O(l) + Pb(s)$

b. $I_2(s) + NaS_2O_3(aq) \longrightarrow Na_2S_2O_4(aq) + NaI(aq)$

Mastering Problems 71

Use the half-reaction method to balance these equations. Add water molecules and hydrogen ions or hydroxide ions as needed.

a. $PbO(s) + NH_3(g) \longrightarrow N_2(g) + H_2O(l) + Pb(s)$

$NH_3(g) \longrightarrow N_2(g) + 3e^-$ oxidation

$PbO(s) + 2e^- \longrightarrow Pb(s)$ reduction

b. $I_2(s) + NaS_2O_3(aq) \longrightarrow Na_2S_2O_4(aq) + NaI(aq)$

$I_2(s) + 2e^- \longrightarrow 2I^-(aq)$ reduction

$2S_2O_3^{2-}(aq) \longrightarrow 2S_2O_4^{2-}(aq) + 2e^-$ oxidation

Mastering Problems 71

Use the half-reaction method to balance these equations. Add water molecules and hydrogen ions or hydroxide ions as needed.

c. $Sn(s) + 2HCl(aq) \longrightarrow SnCl_2(aq) + H_2(g)$

Mastering Problems 71

Use the half-reaction method to balance these equations. Add water molecules and hydrogen ions or hydroxide ions as needed.

c. $Sn(s) + 2HCl(aq) \longrightarrow SnCl_2(aq) + H_2(g)$

$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^-$ oxidation

$2H^+(aq) + 2e^- \longrightarrow H_2(g)$ reduction

Mixed Review 76

Write a balanced ionic redox equation using the following pairs of redox half-reactions.

a. $Fe \longrightarrow Fe^{2+} + 2e^-$

$Te^{2+} + 2e^- \longrightarrow Te$

Mixed Review 76

Write a balanced ionic redox equation using the following pairs of redox half-reactions.

a. $Fe \longrightarrow Fe^{2+} + 2e^-$

$Te^{2+} + 2e^- \longrightarrow Te$

$Fe \longrightarrow Fe^{2+} + \cancel{2e^-}$

$Te^{2-} + \cancel{2e^{2}} \longrightarrow Te$

$Fe + Te^{2+} \longrightarrow Fe^{2+} + Te$

Mixed Review 76

Write a balanced ionic redox equation using the following pairs of redox half-reactions.

b. $IO_4^- + 2e^- \longrightarrow IO_3^-$

$Al \longrightarrow Al^{3+} + 3e^-$ (in acid solution)

Mixed Review 76

Write a balanced ionic redox equation using the following pairs of redox half-reactions.

b. $IO_4^- + 2e^- \longrightarrow IO_3^-$

$Al \longrightarrow Al^{3+} + 3e^-$ (in acid solution)

$3(\overset{+7}{I}O_4^- + 2H^+ + 2e^- \longrightarrow \overset{+5}{I}O_3^- + H_2O)$

$2(\overset{0}{Al} \longrightarrow \overset{+3}{Al}^{3+} + 3e^-)$

$3IO_4^- + 2Al + 6H^+ \longrightarrow 3IO_3^- + 2Al^{3+} + 3H_2O$

Mixed Review 76

Write a balanced ionic redox equation using the following pairs of redox half-reactions.

c. $I_2 + 2e^- \longrightarrow 2I^-$

$N_2O \longrightarrow NO_3^- + 4e^-$ (in acid solution)

Mixed Review 76

Write a balanced ionic redox equation using the following pairs of redox half-reactions.

c. $I_2 + 2e^- \longrightarrow 2I^-$

$N_2O \longrightarrow NO_3^- + 4e^-$ (in acid solution)

$4(I_2 + \cancel{2e^-} \longrightarrow 2I^-)$

$N_2O + 5H_2O \longrightarrow 2NO_3^- + 10H^+ + \cancel{8e^-}$

$4I_2 + N_2O + 5H_2O \longrightarrow 8I^- + 2NO_3^- + 10H^+$

Mixed Review 78

Balance these ionic redox equations by any method.

a. $Sb^{3+} + MnO_4^- \longrightarrow SbO_4^{3-} + Mn^{2+}$ (in acid solution)

Mixed Review 78

Balance these ionic redox equations by any method.

a. $Sb^{3+} + MnO_4^- \longrightarrow SbO_4^{3-} + Mn^{2+}$ (in acid solution)

Oxidation: $5(\overset{+3}{Sb}^{3+} + 4H_2O \longrightarrow \overset{+5}{Sb}\,O_4^{3-} + 8H^+ + 2e^-)$

Reduction: $2(\overset{+7}{Mn}\,O_4^- + 8H^+ + 5e^- \longrightarrow \overset{+2}{Mn}^{2+} + 4H_2O)$

$5Sb^3 + \cancel{20H_2O} + 2MnO_4^- + \cancel{16H^+} + \cancel{10e^-} \longrightarrow 5SbO_4^{3-} + \cancel{40H^+} + 2Mn^{2+} + \cancel{20H_2O} + \cancel{10e^-}$

$5Sb^{3+} + 2MnO_4^- + 12H_2O \longrightarrow 5SbO_4^{3-} + 2Mn^{2+} + 24H^+$

Mixed Review 78

Balance these ionic redox equations by any method.

b. $N_2O + ClO^- \longrightarrow Cl^- + NO_2^-$ (in basic solution)

Mixed Review 78

Balance these ionic redox equations by any method.

b. $N_2O + ClO^- \longrightarrow Cl^- + NO_2^-$ (in basic solution)

Oxidation: $N_2O + 6OH^- \longrightarrow 2NO_2^- + 3H_2O + 4e^-$

Reduction: $H_2O + ClO^- + 2e^- \longrightarrow Cl^- + 2OH^-$

$2H_2O + 2ClO^- + 4e^- \longrightarrow 2Cl^- + 4OH^-$

$N_2O + 6OH^- + 2H_2O + 2ClO^- + 4e^- \longrightarrow 2NO_2^- + 3H_2O + 4e^- + 2Cl^- + 4OH^-$

$N_2O + 2OH^- + 2ClO^- \longrightarrow 2NO_2^- + H_2O + 2Cl^-$

Page 156

4.2 Balancing Redox Equations

Mixed Review 80

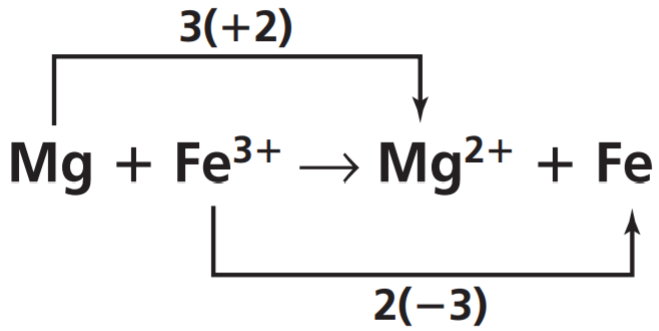
Balance these ionic redox equations by any method.

a. $Mg + Fe^{3+} \longrightarrow Mg^{2+} + Fe$

Mixed Review 80

Balance these ionic redox equations by any method.

a. $Mg + Fe^{3+} \longrightarrow Mg^{2+} + Fe$



$$3Mg + 2Fe^{3+} \longrightarrow 3Mg^{2+} + 2Fe$$

Mixed Review 80

Balance these ionic redox equations by any method.

b. $ClO_3^- + SO_2 \longrightarrow Cl^- + SO_4^{2-}$ (in acid solution)

Mixed Review 80

Balance these ionic redox equations by any method.

b. $ClO_3^- + SO_2 \longrightarrow Cl^- + SO_4^{2-}$ (in acid solution)

$3(\overset{+4}{S}O_2 + 2H_2O \rightarrow \overset{+6}{S}O_4^{2-} + 4H + 2e^-)$

$\overset{+5}{Cl}O_3^- + 6H^+ + 6e^- \rightarrow \overset{-1}{Cl}^- + 3H_2O$

$ClO_3^- + \cancel{6H_2O} + 3SO_2 + \cancel{6H^+} + 6e^- \rightarrow$
$3SO_4^{2-} + \overset{3}{\cancel{12H^+}} + Cl^- + \cancel{3H_2O} + 6e^-$

$3(\overset{+4}{S}O_2 + \overset{6}{2}H_2O \rightarrow \overset{+6}{S}O_4^{2-} + 4H^+ + 2e^-)$

$\overset{+5}{Cl}O_3^- + 6H^+ + 6e^- \rightarrow \overset{-1}{Cl}^- + 3H_2O$

$ClO_3^- + \overset{3}{\cancel{6}}H_2O + 3SO_2 + \cancel{6H^+} + \cancel{6e^-} \rightarrow$
$3SO_4^{2-} + \overset{6}{\cancel{12}}H^+ + Cl^- + \cancel{3H_2O} + \cancel{6e^-}$

$3SO_2 + ClO_3^- + 3H_2O \rightarrow 3SO_4^{2-} + Cl^- + 6H^+$

Page 156

4.2 Balancing Redox Equations

Mixed Review 81

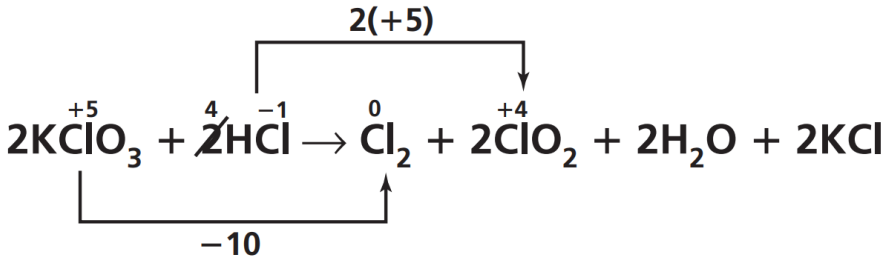
Balance these redox equations by any method.

b. $KClO_3 + HCl \longrightarrow Cl_2 + ClO_2 + H_2O + KCl$

Mixed Review 81

Balance these redox equations by any method.

b. $KClO_3 + HCl \longrightarrow Cl_2 + ClO_2 + H_2O + KCl$



$2KClO_3 + 4HCl \longrightarrow Cl_2 + 2ClO_2 + 2H_2O + 2KCl$

$HCl \longrightarrow ClO_2 + 5e^-$

$2HCl \longrightarrow 2ClO_2 + 10e^-$

$2KClO_3 + 2H^- + 2Cl^- + 2HCl + 10e^- \longrightarrow Cl_2 + 2ClO_2 + 2H_2O + 2K^+ + 2Cl^-$

$2KClO_3 + 4HCl \longrightarrow Cl_2 + 2ClO_2 + 2H_2O + 2KCl$

Page 156

4.2 Balancing Redox Equations

Think Critically 83

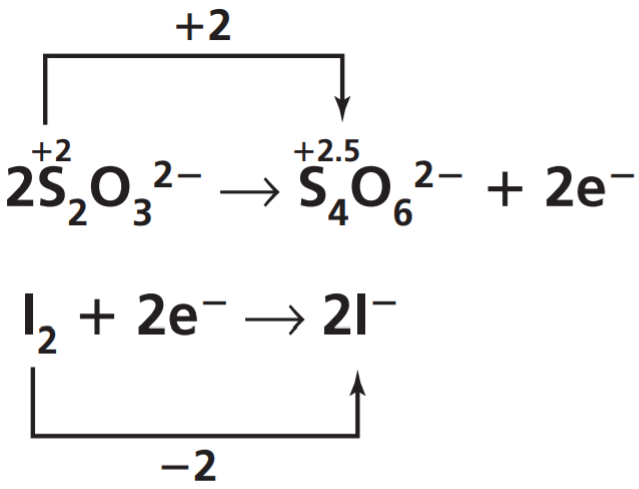
Balance the equation using the half-reaction method.

$$S_2O_3^{2-} + I_2 \rightarrow I^- + S_4O_6^{2-} \text{ (in acidic solution)}$$

Think Critically 83

Balance the equation using the half-reaction method.

$$S_2O_3^{2-} + I_2 \rightarrow I^- + S_4O_6^{2-} \text{ (in acidic solution)}$$



$$2S_2O_3^{2-} + I_2 \rightarrow 2I^- + S_4O_6^{2-} \text{ (in acid solution)}$$

Page 156

4.2 Balancing Redox Equations

Think Critically 85

Potassium permanganate oxidizes chloride ions to chlorine gas. Balance the equation for this redox reaction taking place in acidic solution.

Think Critically 85

Potassium permanganate oxidizes chloride ions to chlorine gas. Balance the equation for this redox reaction taking place in acidic solution.

$$2MnO^{4-}\,(aq) + 10Cl^{-}\,(aq) + 16H^{+}\,(aq) \rightarrow 2Mn^{2+}\,(aq) + 5Cl_2\,(aq) + 8H_2O\,(l)$$

Think Critically 86

In the half-reaction $NO_3^- \rightarrow NH_4^+$, on which side of the equation should electrons be added? Add the correct number of electrons to the side on which they are needed, and rewrite the equation.

Think Critically 86

In the half-reaction $NO_3^- \rightarrow NH_4^+$, on which side of the equation should electrons be added? Add the correct number of electrons to the side on which they are needed, and rewrite the equation.

The oxidation state of N is reduced from +5 to −3; N must gain 8 electrons.

$8e^-$ to the left side; $NO_3^- + 8e^- \rightarrow NH^{4+}$

Think Critically 87

The redox reaction between dichromate ions and iodide ions in acidic solution is shown in Figure 15. Use the half-reaction method to balance the equation for this redox reaction.

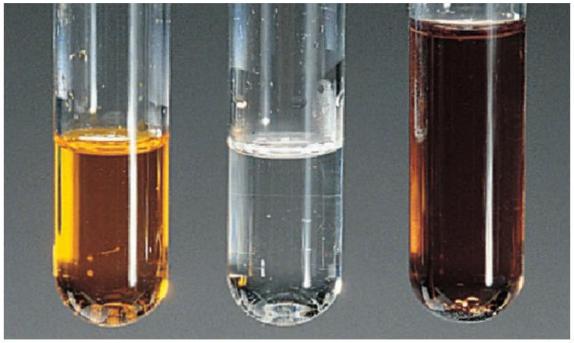


Figure 15

Think Critically 87

The redox reaction between dichromate ions and iodide ions in acidic solution is shown in Figure 15. Use the half-reaction method to balance the equation for this redox reaction.

$2I^- \rightarrow I_2 + 2e^-$

$6I^- \rightarrow 3I_2 + 6e^-$

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

$6I^-(aq) + 14H^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow 3I_2(s) + 7H_2O(l) + 2Cr^{3+}(aq)$

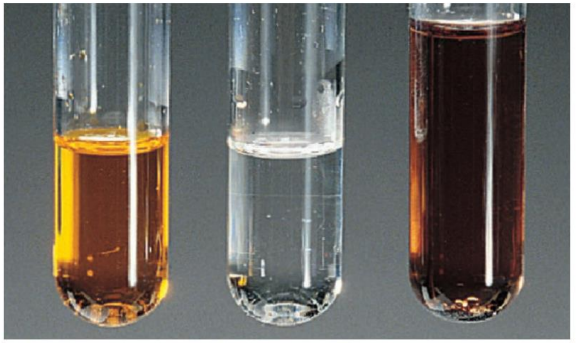


Figure 15

Multiple Choice 2

The reaction between nickel and copper(II) chloride is shown below.

$$Ni(s) + CuCl_2(aq) \rightarrow Cu(s) + NiCl_2(aq)$$

What are the half-reactions for this redox reaction?

A. $Ni \rightarrow Ni^{2+} + 2e^-$, $Cl_2 \rightarrow 2Cl^- + 2e^-$
B. $Ni \rightarrow Ni^{2+} + e^-$, $Cu^+ + e^- \rightarrow Cu$
C. $Ni \rightarrow Ni^{2+} + 2e^-$, $Cu^{2+} + 2e^- \rightarrow Cu$
D. $Ni \rightarrow Ni^{2+} + 2e^-$, $2Cu^+ + 2e^- \rightarrow Cu$

Multiple Choice 2

The reaction between nickel and copper(II) chloride is shown below.

$$Ni(s) + CuCl_2(aq) \rightarrow Cu(s) + NiCl_2(aq)$$

What are the half-reactions for this redox reaction?

C. $Ni \rightarrow Ni^{2+} + 2e^-, Cu^{2+} + 2e^- \rightarrow Cu$

Chapter 4 – Redox Reactions

Resources

- *Redox Reactions, from Glencoe Chemistry: Matter and Change ©2017*

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