Chapter 13 SUMMARY

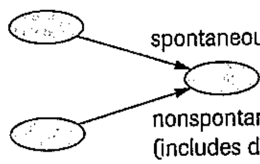
Make a Summary

(Page 604)

1. **Redox**

- redox reactions →
  - **reduction**: gain of electrons half-reaction; decrease in oxidation number; oxidizing agent
  - **oxidation**: loss of electrons half-reaction; increase in oxidation number; reducing agent
- redox table →

OA **RA**



spontaneous

nonspontaneous
(includes disproportionation:
same entity acts as both OA and RA)

- procedures →
  - Writing Half-Reaction Equations (page 567)
  - Five-Step Method for Predicting Redox Reactions (page 578)
  - Predicting Redox Reactions by Constructing Half-Reactions (page 581)
  - Determining Oxidation Numbers (page 585)
  - Balancing Redox Equations using Oxidation Numbers (page 593)

Five-Step Method for Predicting Redox Reactions (page 578)
Predicting Redox Reactions by Constructing Half-Reactions (page 581)
Determining Oxidation Numbers (page 585)
Balancing Redox Equations Using Oxidation Numbers (page 593)

2. (1) Electrochemical reactions are characterized by a transfer of electrons. The entity with the greatest tendency to gain electrons pulls electrons from the entity with the greatest tendency to lose or give up electrons. One entity gains electrons in an electrochemical process while another entity loses electrons.
   (2) Using the experimentally determined redox table, locate the positions of the strongest oxidizing and reducing agents present in the initial mixture. If the strongest oxidizing agent appears above the strongest reducing agent in the table, then a spontaneous redox reaction should occur.
   (3) Redox stoichiometry, like other forms of stoichiometry, uses the same apparatus, procedure, and assumptions (i.e., reactions are stoichiometric). Unlike acid-base titrations, redox titrations are usually self-indicating and no extra indicator is usually required. The procedure of writing redox reactions involves a more complicated set of steps rather than a simple generalization as used previously for predicting chemical reaction equations.

Chapter 13 REVIEW

Part 1

(Pages 605–606)

1. B
2. A
3. D
4. 1, 3, 5, 6
5. B

6. D
7. C
8. C
9. A
10. 3, 1, 1, 3
11. B
12. B
13. A
14. D
15. 65.6

Solutions

15. **OA** **OA** **SOA** **OA** **OA**

$Sn^{2+}(aq)$, $K^{+}(aq)$, $Cr_2O_7^{2-}(aq)$, $H^{+}(aq)$, $Cl^{-}(aq)$, $H_2O(l)$

SRA **RA** **RA**

$$Cr_2O_7^{2-}(aq) + 14\ H^{+}(aq) + 6\ e^{-} \rightarrow 2\ Cr^{3+}(aq) + 7\ H_2O(l)$$
$$3\ [\ Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2\ e^{-}]$$
$$Cr_2O_7^{2-}(aq) + 14\ H^{+}(aq) + 3\ Sn^{2+}(aq) \rightarrow 2\ Cr^{3+}(aq) + 3\ Sn^{4+}(aq) + 7\ H_2O(l)$$

12.7 mL   25.0 mL
c 0.100 mol/L

$$n_{Sn^{2+}} = 25.0\ m\cancel{L} \times \frac{0.100\ \text{mol}}{1\ \cancel{L}} = 2.50\ \text{mmol}$$

$$n_{Cr_2O_7^{2-}} = 2.50\ \text{mmol} \times \frac{1}{3} = 0.833\ \text{mmol}$$

$$[Cr_2O_7^{2-}(aq)] = \frac{0.833\ \cancel{m}\text{mol}}{12.7\ \cancel{m}\text{L}} = 0.0656\ \text{mol/L or } 65.6\ \text{mmol/L}$$

$$\text{or } [Cr_2O_7^{2-}(aq)] = 25.0\ \cancel{m}\ \cancel{L\ Sn^{2+}} \times \frac{0.100\ \cancel{mol\ Sn^{2+}}}{1\ \cancel{L\ Sn^{2+}}} \times \frac{1\ \text{mol}\ Cr_2O_7^{2-}}{3\ \cancel{mol\ Sn^{2+}}} \times \frac{1}{12.7\ \cancel{m}L\ Cr_2O_7^{2-}}$$
$$= 65.6\ \text{mmol/L}$$

On the basis of the evidence, the amount concentration of potassium dichromate solution is 65.6 mmol/L.

Part 2

(Pages 606–609)

16. According to modern theory, a redox reaction involves a transfer of electrons from the reducing agent to the oxidizing agent. The reducing agent loses electrons in an oxidation half-reaction. The oxidizing agent gains those electrons in a reduction half-reaction.
17. (a) In a table of half-reactions, if the oxidizing agent is listed above the reducing agent, then the reaction is predicted to be spontaneous. If the oxidizing agent is listed below the reducing agent, then the reaction is predicted to be non-spontaneous.
    (b) Possible evidence of a spontaneous redox reaction includes the formation of a precipitate or gas, a colour or odour change, or an energy change.
18. (a) Oxidation is defined as a loss of electrons involving an increase in oxidation number.
    (b) Reduction is defined as a gain of electrons involving a decrease in oxidation number.
    (c) In a redox reaction, an electron transfer takes place in which the oxidizing agent gains electrons and the reducing agent loses electrons. An atom/ion in the oxidizing agent decreases in oxidation number and an atom/ion in the reducing agent increases in oxidation number.

19. (a) $HClO_2(aq) + 2\ H^+(aq) + 2\ e^- \rightarrow HClO(aq) + H_2O(l)$(reduction)

(b) $Al(OH)_4^-(aq) + 3\ e^- \rightarrow Al(s) + 4\ OH^-(aq)$(reduction)

(c) $Br^-(aq) + 4\ H_2O(l) \rightarrow BrO_4^-(aq) + 8\ H^+(aq) + 8\ e^-$(oxidation)

(d) $2\ ClO^-(aq) + 2\ H_2O(l) + 2\ e^- \rightarrow Cl_2(aq) + 4\ OH^-(aq)$(reduction)

20. (a) SOA ↓

$Cd^{2+} + 2\ e^- \leftrightarrows Cd(s)$

$Ga^{3+}(aq) + 3\ e^- \leftrightarrows Ga(s)$

$Mn^{2+}(aq) + 2\ e^- \leftrightarrows Mn(s)$

$Ce^{3+}(aq) + 3\ e^- \leftrightarrows Ce(s)$ SRA ↑

(b) The strongest oxidizing agent is Cd^{2+} and the strongest reducing agent is Ce(s).

21. (a)

OA	**OA**
$Cl_2(aq)$,	$H_2O(l)$
	RA

(b)

OA		**OA**
$Sn^{2+}(aq)$,	$NO_3^-(aq)$,	$H_2O(l)$
RA		**RA**

(c)

OA	**OA**	⌒ **OA** ⌒	**OA**
$K^+(aq)$,	$H^+(aq)$,	$IO_3^-(aq)$,	$H_2O(l)$
			RA

22. (a)

SOA	**OA**	⌒ **OA** ⌒	**OA**
$Cl_2(g)$,	$Fe^{2+}(aq)$,	$SO_4^{2-}(aq)$	$H_2O(l)$
	SRA		**RA**

$Cl_2(g) + 2\ e^- \rightarrow 2\ Cl^-(aq)$

$2\ [Fe^{2+}(aq) \rightarrow 2\ Fe^{3+}(aq) + e^-]$

$Cl_2(g) + 2\ Fe^{2+}(aq) \xrightarrow{\text{spont.}} 2\ Cl^-(aq) + 2\ Fe^{3+}(aq)$

(b)

OA		**SOA**	⌒ **OA** ⌒	**OA**
$Ni^{2+}(aq)$,	$NO_3^-(aq)$,	$Sn^{2+}(aq)$,	$SO_4^{2-}(aq)$	$H_2O(l)$
		SRA		**RA**

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

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

$2\ Sn^{2+}(aq) \xrightarrow{\text{nonspont.}} Sn(s) + Sn^{4+}(aq)$

(c)

	OA	⌒ **SOA** ⌒
$Zn(s)$,	$H_2O(l)$	$O_2(g)$
SRA	**RA**	

$O_2(g) + 2\ H_2O(l) + 4\ e^- \rightarrow 4\ OH^-(aq)$

$2\ [Zn(s) \rightarrow Zn^{2+}(aq) + 2\ e^-]$

$O_2(g) + 2\ H_2O(l) + 2\ Zn(s) \xrightarrow{\text{spont.}} 4\ OH^-(aq) + 2\ Zn^{2+}(aq)$

$O_2(g) + 2\ H_2O(l) + 2\ Zn(s) \rightarrow 2\ Zn(OH)_2(s)$

(d)

OA ⌒ **SOA** ⌒	⌒ **OA** ⌒	**OA**		**OA**
$H^+(aq)$,	$SO_4^{2-}(aq)$,	$H_2O(l)$,	$Fe(s)$,	$Na^+(aq)$
		RA	**SRA**	

$SO_4^{2-} + 4\ H^+(aq) + 2\ e^- \rightarrow H_2SO_3(aq) + H_2O(l)$

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

$Fe(s) + SO_4^{2-} + 4\ H^+(aq) \xrightarrow{\text{spont.}} Fe^{2+}(aq) + H_2SO_3(aq) + H_2O(l)$

(e) **OA** **SOA** **OA**

$Na^+(aq)$, $H_2O(l)$, $K^+(aq)$ $SO_3^{2-}(aq)$, $OH^-(aq)$

RA **SRA** **RA**

$$2\ H_2O(l) + 2\ e^- \rightarrow H_2(g) + 2\ OH^-(aq)$$
$$SO_3^{2-}(aq) + 2\ OH^-(aq) \rightarrow SO_4^{2-}(aq) + H_2O(l) + 2\ e^-$$
$$SO_3^{2-}(aq) + H_2O(l) \xrightarrow{\text{spont.}} H_2(g) + SO_4^{2-}(aq)$$

23. **Purpose**

The purpose of this investigation is to test the five-step method for predicting redox reactions.

Problem

What are the products of the reaction of calcium metal with water?

Prediction

According to the five-step method and the redox table, the products of the reaction of calcium metal and water are hydrogen gas and a precipitate of calcium hydroxide.

SOA

$Ca(s)$, $H_2O(l)$

SRA **RA**

$$Ca(s) \rightarrow Ca^{2+}(aq) + 2\ e^-$$
$$2\ H_2O(l) + 2\ e^- \rightarrow H_2(g) + 2\ OH^-(aq)$$
$$Ca(s) + 2\ H_2O(l) \xrightarrow{\text{spont.}} Ca^{2+}(aq) + H_2(g) + 2\ OH^-(aq)$$
$$Ca(s) + 2\ H_2O(l) \xrightarrow{\text{spont.}} Ca(OH)_2(s) + H_2(g)$$

Design

A small quantity of pure calcium metal is added to pure water. Diagnostic tests for hydrogen, the pH of the solution, and a flame test are performed. As a control, the same tests are conducted before the reaction.

Materials

- eye protection
- small sample of $Ca(s)$
- 50 mL pure H_2O
- laboratory burner
- wood splint
- lab apron
- large test tube
- wire loop
- pH paper or meter
- matches

CAUTION: Calcium is a flammable solid, water reactive, and corrosive. Further, it can prove harmful and even fatal if swallowed. It is harmful if inhaled or absorbed through the skin. Contact with the skin can also produce burns. Use protective equipment and extreme caution in performing this experiment.

Procedure

1. Pour 50 mL of distilled water into the test tube.
2. Test using a burning splint, pH paper, and flame test. Record your observations.
3. Place a small piece of $Ca(s)$ into the water and record evidence of the reaction.
4. Place a burning splint near the mouth of the test tube.
5. Once the $Ca(s)$ has stopped reacting, perform diagnostic tests using pH paper and flame test on the liquid.
6. Dispose of the mixture into the sink with lots of running water.

24. (a)

$$Pt(s) + 6\ Cl^-(aq) \rightarrow PtCl_6^{2-}(aq) + 4\ e^- \quad \text{(oxidation)}$$
$$4\ [NO_3^-(aq) + 2\ H^+(aq) + e^- \rightarrow NO_2(g) + H_2O(l) \quad \text{(reduction)}$$
$$Pt(s) + 6\ Cl^-(aq) + 4\ NO_3^-(aq) + 8\ H^+(aq) \rightarrow PtCl_6^{2-}(aq) + 4\ NO_2(g) + 4\ H_2O(l)$$

(b) $2\,[CN^-(aq) + H_2O(l) \rightarrow CNO^-(aq) + 2\,H^+(aq) + 2\,e^-]$ (oxidation)

$ClO_2^-(aq) + 4\,H^+(aq) + 4\,e^- \rightarrow Cl^-(aq) + 2\,H_2O(l)$ (reduction)

$2\,CN^-(aq) + ClO_2^-(aq) \rightarrow 2\,CNO^-(aq) + Cl^-(aq)$

(c) $PH_3(g) \rightarrow \frac{1}{4}\,P_4(s) + 3\,H^+(aq) + 3\,e^-$ (oxidation)

$CrO_4^{2-}(aq) + 4\,H^+(aq) + 3\,e^- \rightarrow Cr(OH)_4^-(aq)$ (reduction)

$PH_3(g) + CrO_4^{2-}(aq) + H^+(aq) \rightarrow \frac{1}{4}\,P_4(s) + Cr(OH)_4^-(aq)$

$PH_3(g) + CrO_4^{2-}(aq) + H_2O(1) \rightarrow \frac{1}{4}\,P_4(s) + Cr(OH)_4^-(aq) + OH^-(aq)$

or $4\,PH_3(g) + 4\,CrO_4^{2-}(aq) + 4\,H_2O(1) \rightarrow P_4(s) + 4\,Cr(OH)_4^-(aq) + 4\,OH^-(aq)$

25. (a) 0 +3 −2 0 +3 −2

$2\,Al(s) + Fe_2O_3(s) \rightarrow 2\,Fe(s) + Al_2O_3(s)$

Al(s) is oxidized and Fe^{3+} in $Fe_2O_3(s)$ is reduced

(b) 0 +1 +3 0

$In(s) + 3\,Tl^+(aq) \rightarrow In^{3+}(aq) + 3\,Tl(s)$

In(s) is oxidized and $Tl^+(aq)$ is reduced

(c) +3 +2 +2 +4

$2\,Cr^{3+}(aq) + Sn^{2+}(aq) \rightarrow 2\,Cr^{2+}(aq) + Sn^{4+}(aq)$

$Cr^{3+}(aq)$ is reduced and $Sn^{2+}(aq)$ is oxidized

(d) 0 −1 −1 0

$Cl_2(aq) + 2\,I^-(aq) \rightarrow 2\,Cl^-(aq) + I_2(aq)$

Cl in $Cl_2(aq)$ is reduced and $I^-(aq)$ is oxidized

(e) +4 −1 0 +2 −1 0

$UCl_4(s) + 2\,Ca(s) \rightarrow 2\,CaCl_2(s) + U(s)$

U^{4+} in $UCl_4(s)$ is reduced and Ca(s) is oxidized

26. (a) 0 +1 −2 0 +4 −2 +1 −2

$C_6H_{12}O_6(s) + 6\,O_2(g) \rightarrow 6\,CO_2(g) + 6\,H_2O(l)$

$4\,e^-/C$ $2\,e^-/O$

$24\,e^-/C_6H_{12}O_6$ $4\,e^-/O_2$

× 1 × 6

(b) +3 +4 −2 +6 −2 0

$2\,Au^{3+}(aq) + 3\,SO_2(g) + 6\,H_2O(l) \rightarrow 3\,SO_4^{2-}(aq) + 2\,Au(s) + 12\,H^+(aq)$

$3\,e^-/Au^{3+}$ $2\,e^-/S$

$3\,e^-/Au^{3+}$ $2\,e^-/SO_2$

× 2 × 3

(c) +5 −2 −2 +1 −2 +4 −2 −1

$2\,BrO_3^-(aq) + C_2H_6O(aq) \rightarrow 2\,CO_2(g) + 2\,Br^-(aq) + 3\,H_2O(l)$

6 e–/Br 6 e–/C

$6\,e^-/BrO_3^-$ $12\,e^-/C_2H_6O$

× 2 × 1

(d) 0 +5 −2 +1 +2 −2

$3\,Ag(s) + NO_3^-(aq) + 4\,H^+(aq) \rightarrow 3\,Ag^+(aq) + NO(g) + 2\,H_2O(l)$

$1\,e^-/Ag$ $3\,e^-/N$

$1\,e^-/Ag$ $3\,e^-/NO_3^-$

× 3 × 1

(e) +1 +5 −2   +4 −2   +1 +6−2   +2 −2

$2\ HNO_3(aq) + 3\ SO_2(g) + 2\ H_2O(l) \rightarrow 3\ H_2SO_4(aq) + 2\ NO(g)$

3 e^-/N   2 e^-/S

3 e^-/HNO_3 2 e^-/SO_2

× 2   × 3

(f) 0   +7 −2   +2 −2 +1   −1

$4\ Zn(s) + BrO_4^-(aq) + 4\ H_2O(l) + 8\ OH^-(aq) \rightarrow 4\ Zn(OH)_4^{2-}(aq) + Br^-(aq)$

2 e^-/Zn   8 e^-/Br

2 e^-/Zn   8 e^-/BrO_4^-

× 4   × 1

27. (a) The three methods are: the five-step method using a redox table of half-reactions; constructing half-reaction equations; and the oxidation number method.
(b) The five-step method using a redox table also predicts the products and the spontaneity.
(c) The oxidation number method and the method of constructing half-reaction equations both require knowledge of the primary products.
(d) If all methods are available, I would prefer the five-step method using a redox table because this method is most efficient and least prone to error because the balanced half-reaction equations are given. *[Other opinions are possible.]*

28. (a) $Zn(s) + Ag_2S(s) \rightarrow ZnS(s) + 2\ Ag(s)$ (*balanced by inspection*)
(b) $Zn \rightarrow Zn^{2+}$: 0 to +2
$2\ Ag^+ \rightarrow 2\ Ag$: +2 to 0
The total increase in oxidation numbers equals the total decrease.
(c) $Ag^+(s) + e^- \rightarrow Ag(s)$ (reduction)
$Zn(s) \rightarrow Zn^{2+}(s) + 2\ e^-$ (oxidation)

29.

	SOA
Mg(s),	$H_2O(l)$
SRA	**RA**

$2\ H_2O(l) + 2\ e^- \rightarrow H_2(g) + 2\ OH^-(aq)$

$Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^-$

$2\ H_2O(l) + Mg(s) \rightarrow H_2(g) + Mg(OH)_2(s)$

2.0 g m

24.31 g/mol   58.33 g/mol

$$n_{Mg} = 2.0\ g \times \frac{1\ mol}{24.31\ g} = 0.082\ mol$$

$$n_{Mg(OH)_2} = 0.082\ mol \times \frac{1}{1} = 0.082\ mol$$

$$m_{Mg(OH)_2} = 0.082\ mol \times \frac{58.33\ g}{1\ mol} = 4.8\ g$$

$$\text{or } m_{Mg(OH)_2} = 2.0\ \cancel{g\ Mg} \times \frac{1\ \cancel{mol\ Mg}}{24.31\ \cancel{g\ Mg}} \times \frac{1\ \cancel{mol\ Mg(OH)_2}}{1\ \cancel{mol\ Mg}} \times \frac{58.33\ g\ Mg(OH)_2}{1\ \cancel{mol\ Mg(OH)_2}}$$

$$= 4.8\ g$$

According to the stoichiometric method, 4.8 g of precipitate forms from a 2.0 g strip of magnesium.

30. **OA** **OA** **SOA** **OA** **OA**

$Fe^{2+}(aq)$, $K^{+}(aq)$, $MnO_4^{-}(aq)$, $H^{+}(aq)$, $H_2O(l)$

SRA **RA**

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

15.0 mL 10.0 mL

7.50 mmol/L c

$$n_{MnO_4^-} = 15.0\text{ mL} \times \frac{7.50\text{ mmol}}{1\text{ L}} = 113\ \mu\text{mol}\quad(\text{or } 0.113\text{ mmol})$$

$$n_{Fe^{2+}} = 133\ \mu\text{mol} \times \frac{5}{1} = 563\ \mu\text{mol}$$

$$[Fe^{2+}(aq)] = \frac{563\ \mu\text{mol}}{10.0\text{ mL}} = 56.3\text{ mmol/L}$$

$$\text{or } [Fe^{2+}(aq)] = 15.0\text{ m}\cancel{\text{L } MnO_4^-} \times \frac{7.50\ \cancel{\text{m}}\ \cancel{\text{mol } MnO_4^-}}{1\ \cancel{\text{L } MnO_4^-}} \times \frac{5\text{ mol } Fe^{2+}}{1\ \cancel{\text{mol } MnO_4^-}} \times \frac{1}{10.0\ \cancel{\text{m}}\text{L } Fe^{2+}}$$

$$= 56.3\text{ mmol/L}$$

According to the evidence and stoichiometric method, the amount concentration of iron(II) ions is 56.3 mmol/L.

31. **Procedure**

1. Clean three small strips of magnesium with steel wool.
2. Add a few millilitres of each unknown solution into separate clean test tubes.
3. Place a strip of magnesium metal into each solution and record any evidence of reaction.
4. For each solution that was unreactive with magnesium, add a few millilitres of the solution to separate clean test tubes.
5. To each of these test tubes, add a few drops of sodium carbonate solution and record any evidence of a reaction.
6. Dispose of all solutions into the waste beaker.

Expected Evidence

For step 3: Two solutions show no change with magnesium, and a dark precipitate forms on the metal in the third solution.

For step 5: One solution shows no change when aqueous sodium carbonate has been added, and one solution produces a white precipitate.

32. (a) Excess copper metal is added to a measured volume of the solution and the silver metal precipitate is collected and weighed.

(b)
$$2\,[Ag^{+}(aq) + e^{-} \rightarrow Ag(s)] \qquad \text{(reduction)}$$
$$Cu(s) \rightarrow Cu^{2+}(aq) + 2\,e^{-} \qquad \text{(oxidation)}$$
$$\overline{2\,Ag^{+}(aq) + Cu(s) \rightarrow 2\,Ag(s) + Cu^{2+}(aq)}$$

According to the redox table, copper metal is a reducing agent that will spontaneously reduce silver ions to silver metal.

33. **Purpose**

The purpose of this investigation is to use redox stoichiometry for a chemical analysis.

Problem

What is the freezing point of a sample of windshield-washer fluid?

Analysis

Standardization of $KMnO_4(aq)$:

OA	SOA	OA	OA		OA	OA
$K^+(aq)$,	$MnO_4^-(aq)$,	$H^+(aq)$,	$Fe^{2+}(aq)$,	$NH_4^+(aq)$,	$SO_4^{2-}(aq)$,	$H_2O(l)$
			SRA			RA

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

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

12.4 mL (trials 2-4) 10.00 mL

c 0.331 mol/L

$$n_{Fe^{2+}} = 10.00\ \text{mL} \times \frac{0.331\ \text{mmol}}{1\ \text{L}} = 3.31\ \text{mmol}$$

$$n_{MnO_4^-} = 3.31\ \text{mmol} \times \frac{1}{5} = 0.662\ \text{mmol}$$

$$[MnO_4^-(aq)] = \frac{0.662\ \text{mmol}}{12.4\ \text{mL}} = 0.0534\ \text{mol/L}$$

$$\text{or } [MnO_4^-(aq)] = 10.00\ \cancel{\text{m}}\ \cancel{\text{L Fe}^{2+}} \times \frac{0.331\ \cancel{\text{mol Fe}^{2+}}}{1\ \cancel{\text{L Fe}^{2+}}} \times \frac{1\ \text{mol MnO}_4^-}{5\ \cancel{\text{mol Fe}^{2+}}} \times \frac{1}{12.4\ \cancel{\text{m}}\text{L MnO}_4^-}$$

$$= 0.0534\ \text{mol/L}$$

Chemical analysis of basic methanol:

$$\overset{-2\ +1\ -2\ +1}{CH_3OH(aq)} + 6\ \overset{+7\ -2}{MnO_4^-(aq)} + 8\ OH^-(aq) \rightarrow \overset{+4\ -2}{CO_3^{2-}(aq)} + 6\ \overset{+6\ -2}{MnO_4^{2-}(aq)} + 6\ H_2O(l)$$

$6\ e^-/C$	$1\ e^-/Mn$
$6\ e^-/CH_3OH$	$1\ e^-/MnO_4^-$
× 1	× 6
10.00 mL	11.7 mL (trials 2 & 3)
c	0.0534 mol/L

$$n_{MnO_4^-} = 11.7\ \text{mL} \times \frac{0.0534\ \text{mol}}{1\ \text{L}} = 0.625\ \text{mmol}$$

$$n_{CH_3OH} = 0.625\ \text{mmol} \times \frac{1}{6} = 0.104\ \text{mmol}$$

$$[CH_3OH(aq)] = \frac{0.104\ \text{mmol}}{10.00\ \text{mL}} = 0.0104\ \text{mol/L}$$

$$\text{or } [CH_3OH(aq)] = 11.7\ \cancel{\text{m}}\ \cancel{\text{L MnO}_4^-} \times \frac{0.0534\ \cancel{\text{mol MnO}_4^-}}{1\ \cancel{\text{L MnO}_4^-}} \times \frac{1\ \text{mol CH}_3\text{OH}}{6\ \cancel{\text{mol MnO}_4^-}} \times \frac{1}{10.00\ \cancel{\text{m}}\text{L CH}_3\text{OH}}$$

$$= 0.0104\ \text{mol/L}$$

Because the windshield-washer fluid had been diluted by a factor of 1000, the amount concentration of the original windshield-washer fluid is 10.4 mol/L, according to the evidence presented and the stoichiometric method. According to the graph shown below based on the evidence in Table 5, the freezing point of a 10.4 mol/L solution of methanol is approximately –33 °C.