

20 Applications of Oxidation/Reduction Titrations

20A AUXILIARY OXIDIZING AND REDUCING REAGENTS

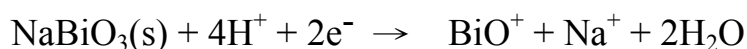
20A-1 Auxiliary Reducing Reagents

Table 20-1 Uses of the Walden Reductor and the Jones Reductor

Walden	Jones
$\text{Ag(s)} + \text{Cl}^- \rightarrow \text{AgCl(s)} + \text{e}^-$	$\text{Zn(Hg)(s)} \rightarrow \text{Zn}^{2+} + \text{Hg} + 2\text{e}^-$
$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$
$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu(s)}$
$\text{H}_2\text{MoO}_4 + 2\text{H}^+ + \text{e}^- \rightarrow \text{MoO}_2^+ + 2\text{H}_2\text{O}$	$\text{H}_2\text{MoO}_4 + 6\text{H}^+ + 3\text{e}^- \rightarrow \text{Mo}^{3+} + 4\text{H}_2\text{O}$
$\text{UO}_2^{2+} + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{U}^{4+} + 2\text{H}_2\text{O}$	$\text{UO}_2^{2+} + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{U}^{4+} + 2\text{H}_2\text{O}$
	$\text{UO}_2^{2+} + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{U}^{3+} + 2\text{H}_2\text{O}$
$\text{V(OH)}_4^+ + 2\text{H}^+ + \text{e}^- \rightarrow \text{VO}^{2+} + 3\text{H}_2\text{O}$	$\text{V(OH)}_4^+ + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{V}^{2+} + 4\text{H}_2\text{O}$
TiO ²⁺ not reduced	$\text{TiO}^{2+} + 2\text{H}^+ + \text{e}^- \rightarrow \text{Ti}^{3+} + 4\text{H}_2\text{O}$
Cr ³⁺ not reduced	$\text{Cr}^{3+} + \text{e}^- \rightarrow \text{Cr}^{2+}$

20A-2 Auxiliary Oxidizing Reagents

Sodium Bismuthate: $\text{Mn(II)} \rightarrow \text{MnO}_4^-$

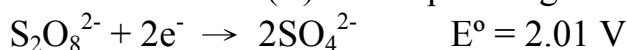


Ammonium Peroxydisulfate, ammonium persulfate, (NH₄)₂S₂O₈

in acidic soln: Cr(III) \rightarrow dichromate

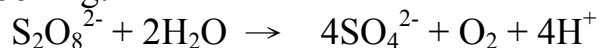
Ce(III) \rightarrow Ce(IV)

Mn(II) \rightarrow permanganate



The oxidations are catalyzed by traces of silver ion.

The excess reagent is readily decomposed by a brief period of boiling:



Sodium Peroxide and Hydrogen Peroxide

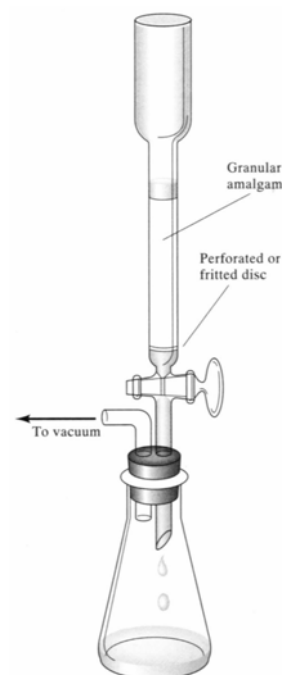
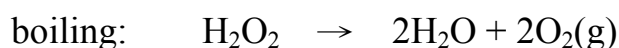
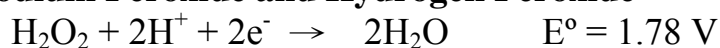


Fig. 20-1
A Jones reductor.

20B APPLING STANDARD REDUCING AGENTS

20B-1 Iron (II) Solutions

iron(II) ammonium sulfate, $\text{Fe(NH}_4)_2(\text{SO}_4)_2 \cdot 6\text{H}_2\text{O}$ (**Mohr's salt**)

iron(II) ethylenediamine sulfate, $\text{FeC}_2\text{H}_4(\text{NH}_3)_2(\text{SO}_4)_2 \cdot 4\text{H}_2\text{O}$ (**Oesper's salt**)

Air-oxidation of iron (II) takes place rapidly in neutral solutions but is inhibited in the presence of acids, with the most stable preparations being about 0.5 M in H₂SO₄.

oxidizing agents ← excess of standard Fe(II) ← standard soln of pot. dichromate or Ce(IV).

Application: organic peroxides, hydroxylamine, Cr(VI), Ce(IV), Mo(VI), nitrate, chlorate, perchlorate and numerous other oxidants.

20B-2 Sodium Thiosulfate

iodine ← thiosulfate

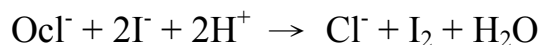


excess KI

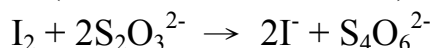
↓ analyte/slightly acidic solution

iodine ← standard solution of Na₂S₂O₃

ex: determination of sod. Hypochlorite in bleaches

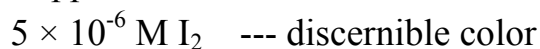


(unmeasured excess KI)



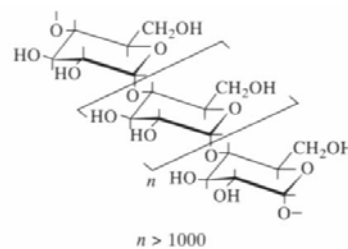
Detecting End Points in Iodine/Thiosulfate Titrations

1. disappearance of the iodine color

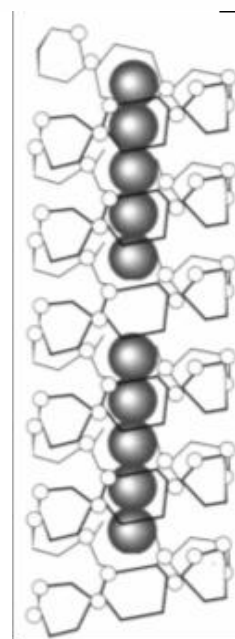


2. starch indicator -- deep blue color

Starch undergoes decomposition in solution with high I₂ concentration. In titration of excess I₂ with Na₂S₂O₃, addition of the indicator must be deferred until most of the I₂ has been reduced.



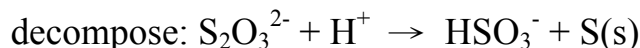
(a)



(b)

Fig. 20-2 Thousands of glucose molecules polymerize to form huge molecules of β-amylose as shown in (a). Molecules of β-amylose tend to assume a helical structure. The iodine species I₃⁻ as shown in (b) is incorporated into the amylose helix.

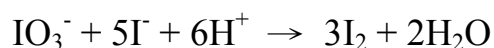
The Stability of Sodium Thiosulfate Solutions



pH, microorganisms, concentration of the solution, presence of Cu(II) ion and exposure to sunlight.

Standardizing Thiosulfate Solutions

primary standard: pot. iodate/excess KI (pot. dichromate, pot. bromate, pot. hydrogen iodate, pot. ferricyanide and metallic copper)/ excess KI.



↑
thiosulfate



Ex. 20-1 A solution of Sod. thiosulfate was standardized by dissolving 0.1210 g KIO_3 (214.00 g/mol) in water, adding a large excess of KI, and acidifying with HCl. The liberated I_2 required 41.64 mL of the thiosulfate soln to decolorize the blue starch/iodine complex. Calculate the molarity of the $\text{Na}_2\text{S}_2\text{O}_3$.

$$\text{amount Na}_2\text{S}_2\text{O}_3 = 0.1210\text{g KIO}_3 \times \frac{1\text{ mol}}{0.21400\text{ g}} \times 6 = 3.3925\text{ mmol}$$

$$C_{\text{Na}_2\text{S}_2\text{O}_3} = \frac{3.3925\text{ mmol}}{41.64\text{ mL}} = 0.08147\text{ M} \quad \text{or} \quad \frac{\frac{121\text{ mg}}{214\text{ mg/mmol}} \times 3 \times 2}{41.64\text{ mL}} = 0.08147\text{ M}$$

Tab 20-2 Applications of Sodium Thiosulfate as Reductant

Analyte	Half-Reaction	Special Condition
IO_4^-	$\text{IO}_4^- + 8\text{H}^+ + 7\text{e}^- \rightarrow \frac{1}{2}\text{I}_2 + 4\text{H}_2\text{O}$	Acid solution
	$\text{IO}_4^- + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{IO}_3^- + \text{H}_2\text{O}$	Neutral solution
IO_3^-	$\text{IO}_3^- + 6\text{H}^+ + 5\text{e}^- \rightarrow \frac{1}{2}\text{I}_2 + 3\text{H}_2\text{O}$	Strong acid
BrO_3^- , ClO_3^-	$\text{XO}_3^- + 6\text{H}^+ + 6\text{e}^- \rightarrow \text{X}^- + 3\text{H}_2\text{O}$	Strong acid
Br_2 , Cl_2	$\text{X}_2 + 2\text{I}^- \rightarrow \text{I}_2 + 2\text{X}^-$	
NO_2^-	$\text{HNO}_2 + \text{H}^+ + \text{e}^- \rightarrow \text{NO}(\text{g}) + \text{H}_2\text{O}$	
Cu^{2+}	$\text{Cu}^{2+} + \text{I}^- + \text{e}^- \rightarrow \text{CuI}(\text{s})$	
O_2	$\text{O}_2 + 4\text{Mn}(\text{OH})_2(\text{s}) + 2\text{H}_2\text{O} \rightarrow 4\text{Mn}(\text{OH})_3(\text{s})$	Basic solution
	$\text{Mn}(\text{OH})_3(\text{s}) + 3\text{H}^+ + \text{e}^- \rightarrow \text{Mn}^{2+} + 3\text{H}_2\text{O}$	Acidic solution
O_3	$\text{O}_3(\text{g}) + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{O}_2(\text{g}) + \text{H}_2\text{O}$	
Organic peroxide	$\text{ROOH} + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{ROH} + \text{H}_2\text{O}$	

20C APPLYING STANDARD OXIDING AGENTS

Table 20-3 Some common oxidants used as standard solutions

Reagent and Formula	Reduction product	Standard Potential, V	Standardized with	Indicator*	Stability#
KMnO_4	Mn^{2+}	1.51†	$\text{Na}_2\text{C}_2\text{O}_4$, Fe, As_2O_3	MnO_4^-	(b)
KBrO_3	Br^-	1.44†	KBrO_3	(1)	(a)
Ce(IV), Ce^{4+}	Ce^{3+}	1.44†	$\text{Na}_2\text{C}_2\text{O}_4$, Fe, As_2O_3	(2)	(a)
$\text{K}_2\text{Cr}_2\text{O}_7$	Cr^{3+}	1.33†	$\text{K}_2\text{Cr}_2\text{O}_7$, Fe	(3)	(a)
I_2	I^-	0.536†	$\text{BaS}_2\text{O}_3 \cdot \text{H}_2\text{O}$, $\text{Na}_2\text{S}_2\text{O}_3$	starch	(c)

* (1) α -Naphthoflavone; (2) 1,10-phenanthroline iron(II) complex (ferroin); (3) diphenylamine sulfonic acid.

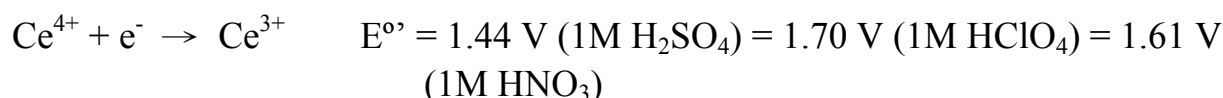
(a) **Indefinitely stable**;

(b) **moderately stable**, requires periodic standardization;

(c) **somewhat unstable**, require frequent standardization.

† E° in H_2SO_4 .

20C-1 The Strong Oxidants-Potassium Permanganate and Cerium(IV)



Comparison of the Two Reagents

	Ce^{4+}	MnO_4^-
in sulfuric acid	stable	decompose slowly
	not oxidize Cl^- --	oxidize Cl^-
HCl soln of analyte	can be used	cannot be used
primary-standard-grade salt	available	
self-indicator	no	color of MnO_4^-
cost (1L 0.02 M soln)	\$ 2.20 (4.40)	\$ 0.08
in < 0.1 M strong acid	tendency to form ppt	

Detecting the End Points

indicators: KMnO_4 solution -- intense purple color

diphenylamine sulfonic acid

1, 10-phenanthroline complex of Fe(II)



rate: slow \rightarrow end point fades only gradually over 30s.

in Ce(IV) titration:

indicator: Fe(II) complex of 1,10-phenanthroline or one of its substitute derivatives (Table 20-3)



Ferrous complex Ferric complex

(Ferriin) red (Ferriin) weak blue

The Preparation and Stability of Standard Solutions

KMnO_4 soln: not entirely stable



decomposition reaction is slow -- catalyzed by **light**, heat, acids, bases, Mn(II) and MnO_2 .

Ex. 20-2 Described how you would prepare 2.0 L of an approximately 0.010 M soln of KMnO_4 (158.03 g/mol).

$$\text{KMnO}_4 \text{ needed} = 2.0 \text{ L} \times 0.010 \text{ M} \times 158.03 \text{ g/mol} = 3.16 \text{ g}$$

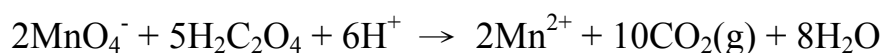
Dissolve about 3.2 g of KMnO_4 in a little water. After solution is complete add water to bring the volume to about 2.0 L. Heat the solution to boiling for a brief period, and let stand until it is cool. Filter through a glass-filtering crucible and stored in a clean dark bottle.

Analytically Useful Cerium(IV) Compounds

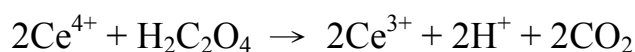
Name	Formula	Molar Mass
Cerium(IV) ammonium nitrate	$\text{Ce}(\text{NO}_3)_4 \cdot 2\text{NH}_4\text{NO}_3$	548.2
Cerium(IV) ammonium sulfate	$\text{Ce}(\text{SO}_4)_2 \cdot 2(\text{NH}_4)_2\text{SO}_4 \cdot 2\text{H}_2\text{O}$	632.6
Cerium(IV) hydroxide	$\text{Ce}(\text{OH})_4$	208.1
Cerium(IV) hydrogen sulfate	$\text{Ce}(\text{HSO}_4)_4$	528.4

Primary Standards

Sodium Oxalate.



Mn(II) as a catalyst (autocatalysis)



Ex. 20-3 You wish to standardize the soln in Ex.20-2 against pure $\text{Na}_2\text{C}_2\text{O}_4$ (134.00 g/mol). If you want to use between 30 and 45 mL of the reagent for the standardization, what range of masses of the primary standard should you weigh out?

for a 30-mL titration:

$$\text{amount KMnO}_4 = 30 \text{ mL} \times 0.010 \text{ M} = 0.30 \text{ mmol}$$

$$\text{mass Na}_2\text{C}_2\text{O}_4 = 0.30 \text{ mmol} \times 5/2 \times 0.134 = \underline{0.101 \text{ g}}$$

for a 45-mL titration:

$$\text{mass Na}_2\text{C}_2\text{O}_4 = 45 \times 0.010 \times 5/2 \times 0.134 = \underline{0.151 \text{ g}}$$

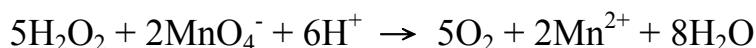
Ex. 20-4 A 0.1278-g sample of primary-standard $\text{Na}_2\text{C}_2\text{O}_4$ required exactly 33.31 mL of the KMnO_4 solution in Ex. 20-2 to reach the end point. What was the molarity of the KMnO_4 reagent?

$$\text{amount Na}_2\text{C}_2\text{O}_4 = 0.1278 \text{ g} \times 1 \text{ mmol}/0.134 \text{ g} = 0.95373 \text{ mmol}$$

$$C_{\text{KMnO}_4} = 0.95373 \text{ mmol} \times (2/5) \times (1/33.31) = 0.01145 \text{ M}$$

Using Potassium Permanganate and Cerium(IV) Solutions: Table 20-5

Ex. 20-5 Aqueous solution containing approximately 3% (w/w) H_2O_2 are sold in drug stores as a disinfectant. propose a method for determining the peroxide content of such a preparation using the standard soln described in Exs.20-3 and 4. Assume that you wish to use between 35 and 45 mL of the reagent for a titration.



35 - 45 mL reagent:

$$\text{amount KMnO}_4 = (35 \sim 45) \text{ mL} \times 0.01145 \text{ M} = 0.401 \sim 0.515 \text{ mmol}$$

$$\text{amount H}_2\text{O}_2 = (0.401 \sim 0.515) \text{ mmol} \times (5/2) = 1.00 \sim 1.29 \text{ mmol}$$

$$\text{mass sample} = (1.00 \sim 1.29) \times 0.03401 \times (100/3) = 1.1 \sim 1.5 \text{ g}$$

Thus we could weigh out from 1.1 to 1.5 g samples. These should be diluted to perhaps 75 to 100 mL with water and made slightly acidic with dilute H_2SO_4 before titration.

20C-2 Potassium Dichromate



orange

green

in 1 M HCl or H_2SO_4 $E'^\circ = 1.0 \sim 1.1 \text{ V}$

Advantages: stable, can be boiled without decomposition and do not react with HCl, primary-standard reagent is available and at a modest cost.

Disadvantage: lower electrode potential and the slowness reaction.

Preparing Dichromate Solutions

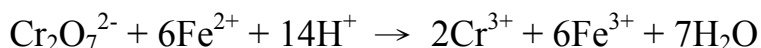
reagent-grade $\text{K}_2\text{Cr}_2\text{O}_7$ dried at 150°C to 200°C before being weighed

indicator: diphenylamine sulfonic acid,

violet (oxidized) \rightarrow colorless (reduced)

Applying Potassium Dichromate Solutions

1. titration of Fe(II): in moderate conc. of HCl

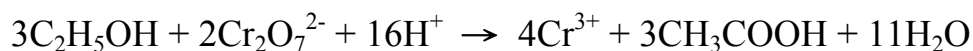


2. indirect determination of oxidizing agents (nitrate, chlorate, permanganate, dichromate and organic peroxides):

analyte/acidic solution + measured excess Fe(II) \rightarrow

back-titrated excess Fe(II)

Ex. 20-6 A 5.00-mL sample of brandy was diluted to 1.000 L in a volumetric flask. The ethanol (C_2H_5OH) in a 25.00-mL aliquot of the diluted soln was distilled into 50.00 mL of 0.02000 M $K_2Cr_2O_7$, and oxidized to acetic acid with heating.



After cooling, 20.00 mL of 0.1253 M Fe^{2+} were pipetted into the flask. The excess Fe^{2+} was then titrated with 7.46 mL of the standard $K_2Cr_2O_7$ to a diphenylamine sulfonic acid end point. Calculate the percent (w/v) C_2H_5OH (46.07 g/mol) in the brandy.

$$\text{amount } K_2Cr_2O_7 = (50.00 + 7.46) \text{ mL} \times 0.02000 = 1.1492 \text{ mmol}$$

$$K_2Cr_2O_7 \text{ consumed by } Fe^{2+} = 20.00 \times 0.1253 \times 1/6 = 0.41767 \text{ mmol}$$

$$K_2Cr_2O_7 \text{ consumed by } C_2H_5OH = 1.1492 - 0.41767 = 0.73153 \text{ mmol}$$

$$\text{mass } C_2H_5OH = 0.73153 \times (3/2) \times 0.04607 = 0.050552 \text{ g}$$

$$\text{percent } C_2H_5OH = 0.050552 / (5.00 \times 25.00 / 1000) \times 100 \% = 40.44 \%$$

20C-3 Iodine

weak oxidizing agents: determination of strong reductants



advantages: selectivity, sensitive and reversible indicator

disadvantage: lack stability

Properties of Iodine Solutions

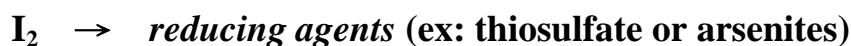


lack stability: volatility of iodine, slowly attacks most organic materials, air-oxidation of iodide ion (\uparrow conc.). $4I^- + O_2(g) + 4H^+ \rightarrow 2I_2 + 2H_2O$

Standardizing and Applying Iodine Solutions

Standardization: anhydrous Na thiosulfate or Ba thiosulfate

Iodimetry: (direct method)

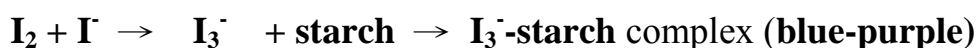


Iodometry: (indirect method)

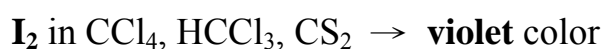


Indicator:

1. Starch indicator solution



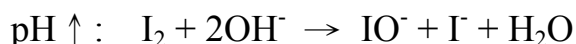
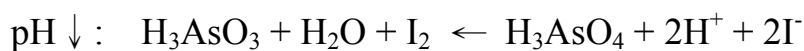
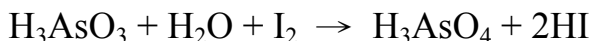
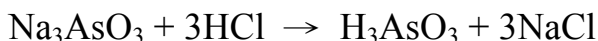
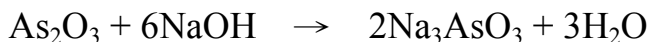
2. CCl_4 , $HCCl_3$, CS_2



Preparation of 0.1 N Iodine solution

12.7 g I₂ + 40 g KI/20 mL H₂O → adding H₂O to 1 L

Standardization- **Primary standard: Arsenic (III) oxide, As₂O₃**



strong oxidizing agent

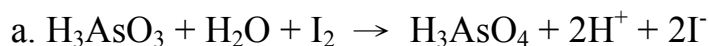
滴定中加入 NaHCO₃ [pH: 7~8]



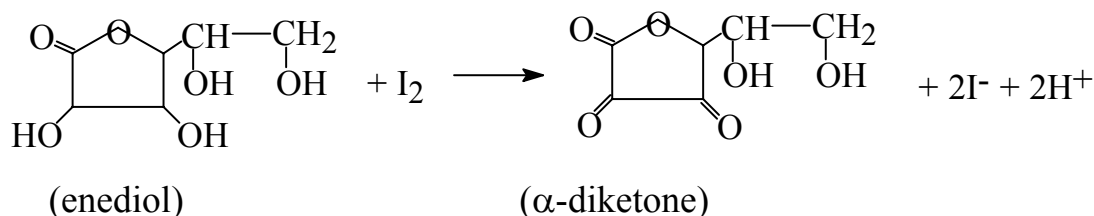
Calculation:

$\text{N of Iodine} = \frac{\frac{\text{mg As}_2\text{O}_3}{197.8/4}}{\text{mL Iodine}}$
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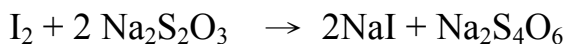
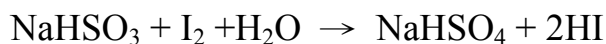
(1) Direct Iodimetric Titration



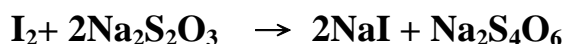
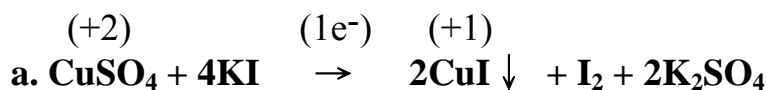
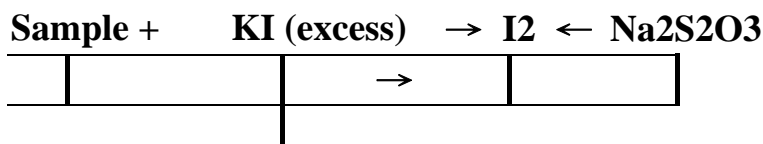
b. Assay of Ascorbic Acid (Vit C)



(2) Residual Titration (I₂- Na₂S₂O₃)



(3) Iodometry:



b. Assay of sodium hypochlorite solution (NaOCl)

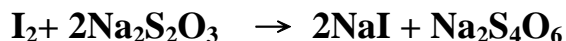
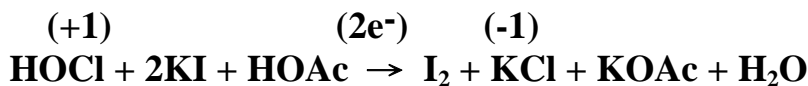
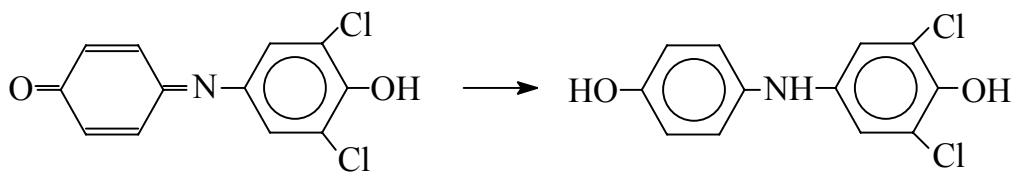


Table 20-6 Some Applications of Iodine Solutions

Analyte	Half-Reaction
As	$\text{H}_3\text{AsO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{AsO}_4 + 2\text{H}^+ + 2e^-$
Sb	$\text{H}_3\text{SbO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{SbO}_4 + 2\text{H}^+ + 2e^-$
Sn	$\text{Sn}^{2+} \rightarrow \text{Sn}^{4+} + 2e^-$
H ₂ S	$\text{H}_2\text{S} \rightarrow \text{S(s)} + 2\text{H}^+ + 2e^-$
SO ₂	$\text{SO}_3^{2-} + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} + 2\text{H}^+ + 2e^-$
S ₂ O ₃ ²⁻	$2 \text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4\text{O}_6^{2-} + 2e^-$
N ₂ H ₄	$\text{N}_2\text{H}_4 \rightarrow \text{N}_2(\text{g}) + 4\text{H}^+ + 4e^-$
Ascorbic acid	$\text{C}_6\text{H}_8\text{O}_6 \rightarrow \text{C}_6\text{H}_6\text{O}_6 + 2\text{H}^+ + 2e^-$

**** Dichloroindophenol Titration**

--Determination of Ascorbic acid Preparation



Blue in basic sol'n

colorless

Pink in acid sol'n

End point: pink (self-indicator)

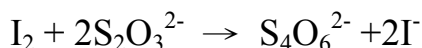
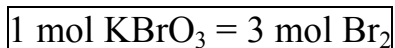
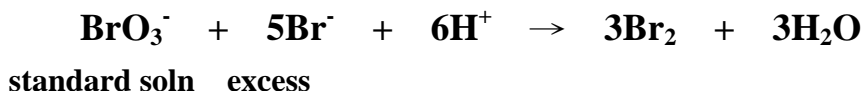
Vit C titration in metaphosphoric acid and acetic acid sol'n

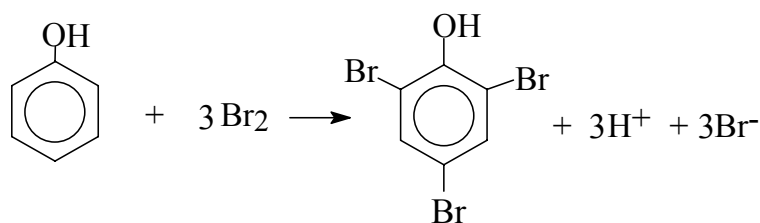
20C-4 Potassium Bromate as a Source of Bromine

Primary-standard KBrO₃ is available, stable

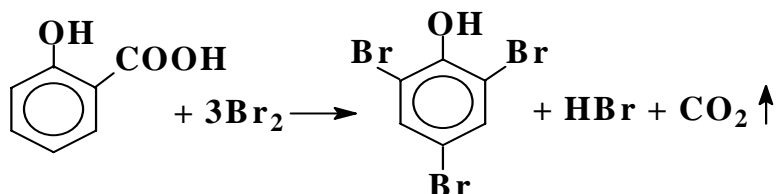
standard 0.1 N Bromine sol'n (Koppeschaar's sol'n): (3 g KBrO₃ + 15 g KBr)/1 L H₂O

Assay of sample: aniline, phenol, salicylic acid, resorcinol etc.





(ppt) soluble in CHCl_3

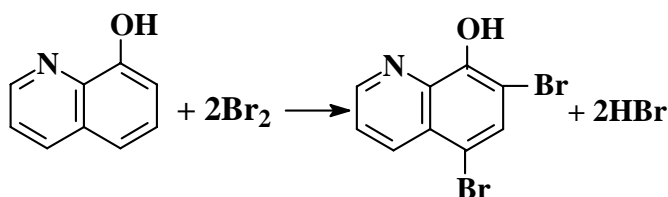


Substitution Reactions

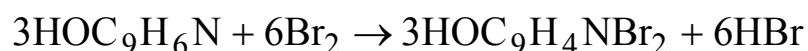
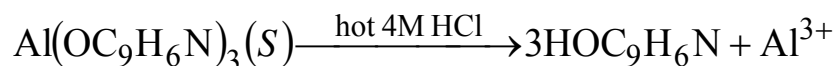
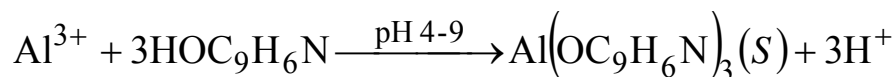
halogen substitution: replacement of H in an aromatic ring by a halogen.

determination of aromatic compound that contain strong ortho-para-directing groups, particularly amines and phenols.

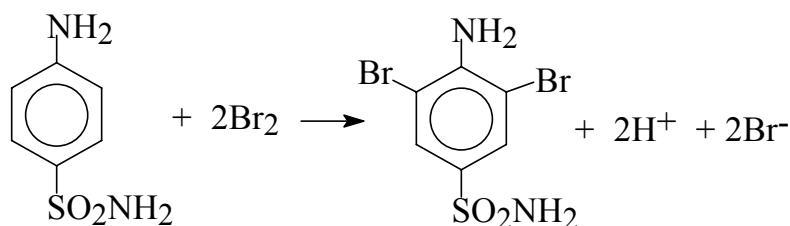
Ex: 1. Determination of 8-hydroxyquinoline

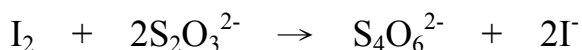


2. Determination of aluminum



Ex. 20-7 A 0.2891-g sample of an antibiotic powder containing sulfanilamide was dissolved in HCl and the solution diluted to 100.0 mL. A 20.00-mL aliquot was transferred to a stoppered flask and 25.00 mL of 0.01767 M KBrO_3 added. About 10 g of KBr was added to form Br_2 , which brominated the sulfanilamide in the sample. After 10 min, an excess of KI was added and the liberated iodine titrated with 12.92 mL of 0.1215 M sodium thiosulfate. The reaction are





Calculate the % $\text{NH}_2\text{C}_6\text{H}_4\text{SO}_2\text{NH}_2$ (172.21 g/mol) in the powder.

total amount $\text{Br}_2 = 25.00 \text{ mL} \times 0.01767 \text{ M} \times 3 = 1.32525 \text{ mmol}$

amount excess $\text{Br}_2 = \text{amount I}_2 = 12.92 \text{ mL} \times 0.1215 \text{ M} \times (1/2) = 0.78489 \text{ mmol}$

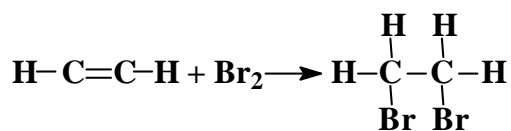
The amount of Br_2 consumed by the sample = $1.32525 - 0.78489 = 0.54036 \text{ mmol}$

mass analyte = $0.54036 \times (1/2) \times 0.17221 = 0.046528 \text{ g}$

$$\% \text{ analyte} = \frac{0.046528}{0.2891 \times \frac{20.00 \text{ mL}}{100 \text{ mL}}} \times 100\% = 80.47\%$$

$$\text{or } \frac{(0.01767 \times 3 \times 2 \times 25 - 0.1215 \times 12.92) \times \frac{172.21}{4} \times 5}{289.1} \times 100\% = 80.47\%$$

Addition Reactions

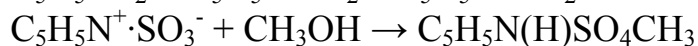
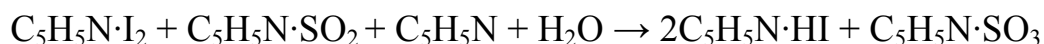


20C-5 Determining Water with the Karl Fischer Reagent

Karl Fischer Reagent: I_2 , SO_2 , organic base such as pyridine ($\text{C}_5\text{H}_5\text{N}$) or imidazole / CH_3OH or low-molecular-mass alcohol

In aprotic solvent: $\text{I}_2 + \text{SO}_2 + 2\text{H}_2\text{O} \rightarrow 2\text{HI} + \text{H}_2\text{SO}_4$ $2 \text{ mol H}_2\text{O} \rightarrow 1 \text{ mol I}_2$

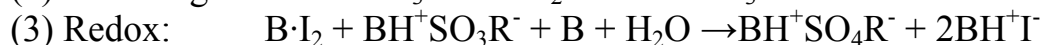
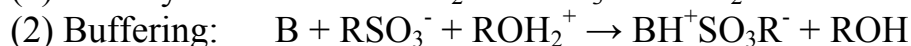
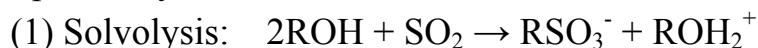
Classical chemistry: use anhydrous methanol as solvent, and excess pyridine



1 mol $\text{H}_2\text{O} \rightarrow 1 \text{ mol I}_2$, 1 mol SO_2 , 3 mol $\text{C}_5\text{H}_5\text{N}$

Pyridine-free chemistry

Replaced by other amines: imidazole



1 mol $\text{H}_2\text{O} \rightarrow 1 \text{ mol I}_2$

Interfering reactions