

BC Science Chemistry 12
Chapter 6 – Oxidation – Reduction and Its Applications Answer Key
September 20, 2012

6.1 Oxidation – Reduction

Warm Up

Name	Formula	Charge on Metal Ion
iron(III) chloride	FeCl ₃	+3
iron(II) acetate	Fe(CH ₃ COO) ₂	+2
manganese(II) bromate	Mn(BrO ₃) ₂	+2
manganese(IV) oxide	MnO ₂	+4
potassium permanganate	KMnO ₄	+7

Practice Problems:

- H₂O H = +1 O = -2
- Cs₂O₂ Cs = +1 O = -1
- CO₃²⁻ C = +4 O = -2
- Na₂Cr₂O₇ Na = +1 Cr = +6 O = -2
- BaH₂ Ba = +2 H = -1
- NH₄⁺ N = -3 H = +1
- S₈ S = 0
- Al₂(SO₄)₃ Al = +3 S = +6 O = -2

Quick Check:

- 6 e-
- nothing (no e- transferred)

Quick Check

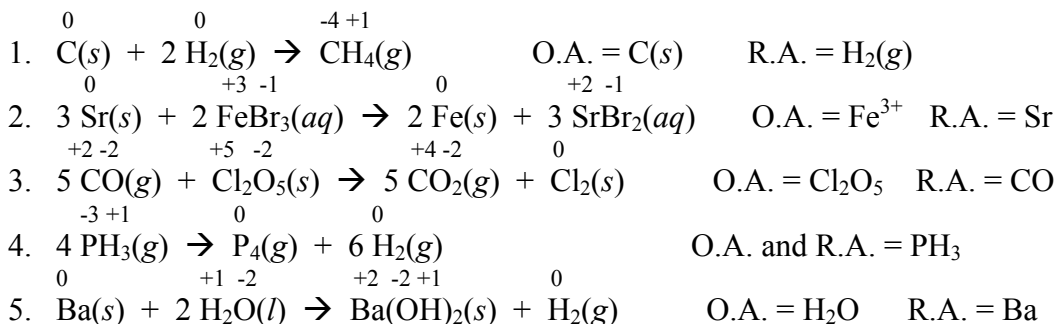
- You – oxidized (LEO)
Partner - reduced (GER)
- Which partner acted as an “oxidizing agent”? my partner (they got reduced)

A “reducing agent”? **me (I got oxidized)**

3. What happened to your “oxidation state”? **Increased (I was oxidized)**

What about your partner? **Reduced – got smaller (they were reduced)**

Practice Problems: Agents



6.1 Activity

Characteristic	Oxidizing Agent	Reducing Agent
Causes another species to be...	oxidized	reduced
Is itself _____ during reaction.	reduced	oxidized
Its oxidation state is __ during reaction.	decreased	increased
Causes electrons to be...	lost	gained
_____ electrons.	gains	loses
Belongs to _____ family.	halogen	alkali METAL
Has a _____ electronegativity.	high	low

6.1 Review Questions

1. 1. Elements that get oxidized (act as reducing agents) form (a) + ions when they react. This means reducing agents are generally (b) metals. Reducing agents may also be (c) negatively charged ions. The most active reducing agents likely belong to the (d) alkali metal family on the periodic table.

The most active oxidizing agents must belong to the (e) halogen family.

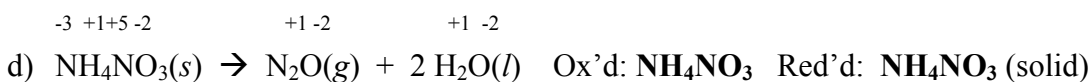
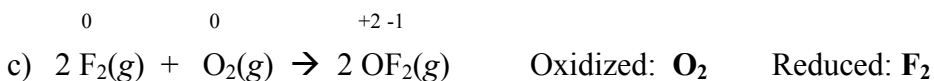
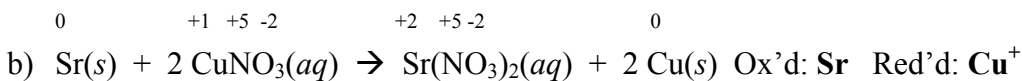
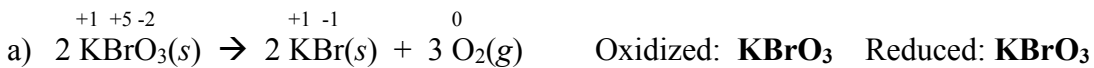
2. CaI₂ b) OF₂ c) C₆H₁₂O₆ d) RbB₂O₂ e) S₂O₃²⁻ f) BeH₂ g) BrO⁻ h) Cl₂
 +2 +2 0 -1 +2 -1 +1 0

3. a) **A species that gets reduced/causes other species to be oxidized/gains e-**

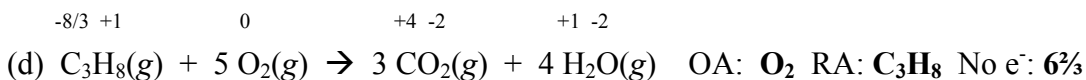
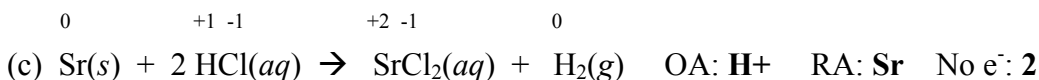
- b) **A species that gets oxidized/causes other species to be reduced/loses e-**

c) **Increased electronegativity makes a stronger O.A. and a weaker R.A.**

4. For each of the following reactions, indicate the species being oxidized and reduced and show the oxidation states above their symbols.



5. Determine the oxidizing and reducing agent in each of the following reactions. Then indicate the number of electrons transferred by one atom of the reducing agent.



6. a) Write a balanced redox equation (in net ionic form) to show what has occurred in the beaker over time: $\mathbf{Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)}$

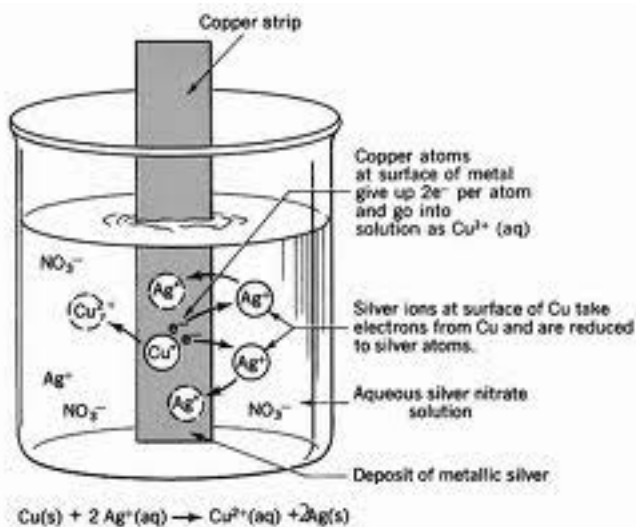
b) What is the oxidizing agent? $\mathbf{Cu^{2+}(aq)}$ The reducing agent? $\mathbf{Fe(s)}$

c) How many electrons were transferred in the equation? $\mathbf{2e^-}$

7. Give the oxidation state of the underlined element in each species:

a) $\underline{\mathbf{P}}^{3-}$ $\mathbf{-3}$ b) $(\text{NH}_4)_2\underline{\mathbf{Zr}}(\text{SO}_4)_3$ $\mathbf{+4}$ c) $\text{Na}_2\underline{\mathbf{C}}_2\text{O}_4$ $\mathbf{+3}$ d) $\underline{\mathbf{N}}_2\text{H}_5\text{Cl}$ $\mathbf{-2}$ e) $\underline{\mathbf{Mn}}\text{O}_4^{2-}$ $\mathbf{+6}$

8.



a) Write a balanced equation to represent the redox reaction occurring in the beaker. (see above)

b) Oxidized? **Cu metal** Reduced? **Ag⁺ ion**

c) Reducing agent? **Cu metal** Oxidizing agent? **Ag⁺ ion**

d) **2e⁻**

a) **Alkali metal family (IA)**

b) **Halogens (VIIA) or (17)**

9. a) **ClO₄⁻, ClO₃⁻, ClO₂⁻**

b) **Cl₂, Cl⁻**

6.2 Balancing Oxidation-Reduction Equations

Warm Up

Does this equation appear to be balanced? yes or no

Which species is oxidized and consequently acts as a reducing agent? Zn

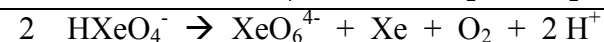
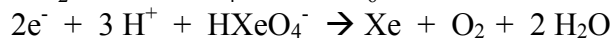
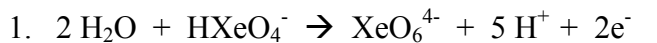
Which species is reduced and consequently acts as an oxidizing agent? Cu⁺

How many electrons are lost by each atom of reducing agent being oxidized? 2

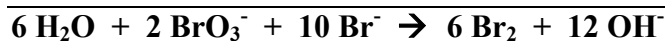
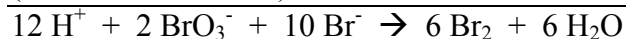
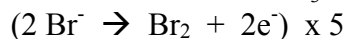
How many electrons are gained by each ion of the oxidizing agent as it is reduced? 1

Rewrite the equation in its balanced form: **Zn(s) + 2 Cu⁺(aq) → Zn²⁺(aq) + 2 Cu(s)**

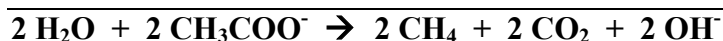
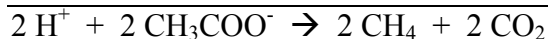
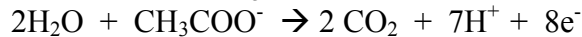
Practice Problems – Disproportionation



Species	Reactants	Products
Hydrogen	4	4
Xenon	2	2
Oxygen	10	10
Charge	-4	-4

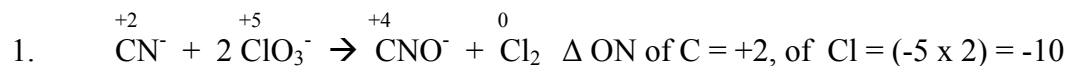


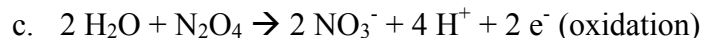
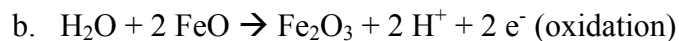
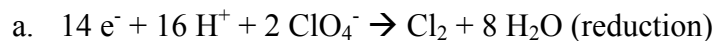
Species	Reactants	Products
Hydrogen	12	12
Bromine	12	12
Oxygen	12	12
Charge	-12	-12



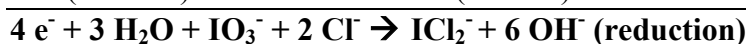
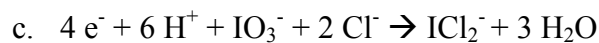
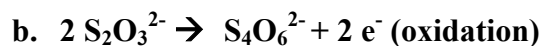
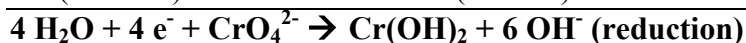
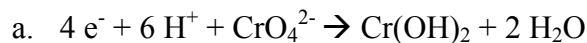
Species	Reactants	Products
Hydrogen	10	10
Carbon	4	4
Oxygen	6	6
Charge	-2	-2

6.2 Activity Key

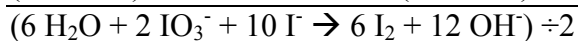
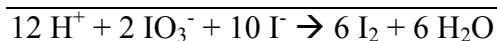
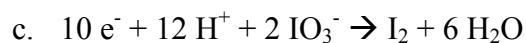
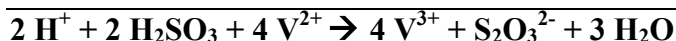
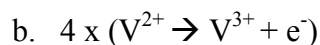
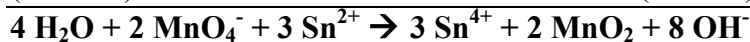
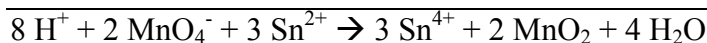
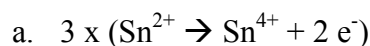


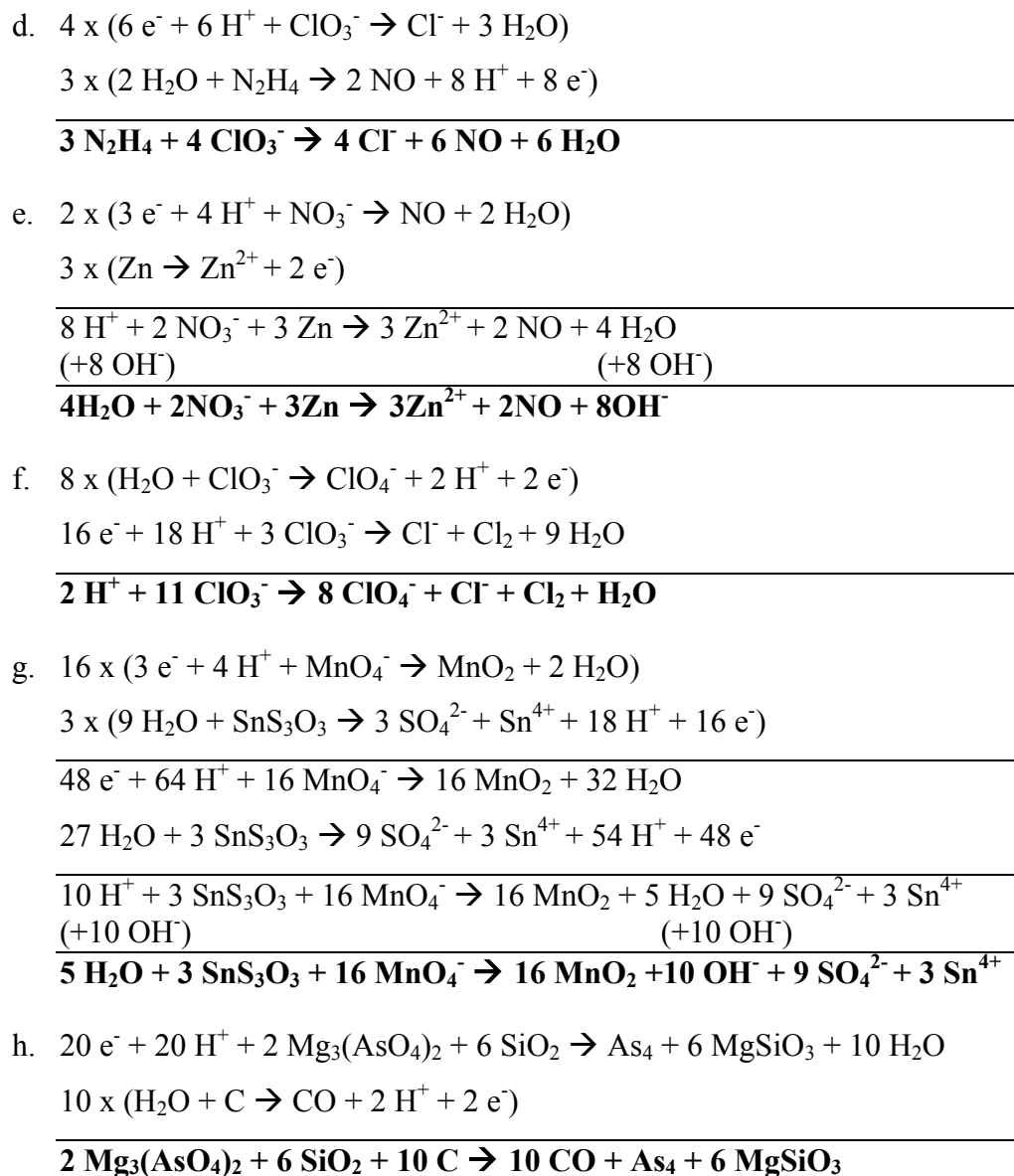


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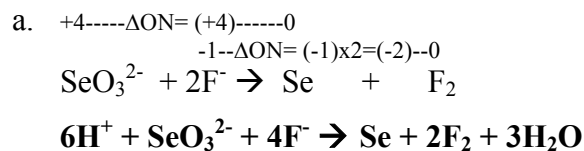


5)





6) Balance



6.3 Using the Standard Reduction Potential (SRP) Table to Predict Redox Reactions

Warm Up

1. For example: both are lists of equilibria, both contain numerical values, both show one type of chemical converting to another (both show only one type of reaction), reactants get stronger as you go up on the left side of both, products get stronger as you go down on the right side of both.
2. reductions; oxidizing agents
3. alkali metals
4. fluorine; halogens

Practice Problems – Determining Whether a Spontaneous Redox Reaction Will Occur

1. yes; $2\text{I}^- + \text{Br}_2 \rightarrow 2\text{Br}^- + \text{I}_2$
2. no; both are oxidizing agents
3. yes; $2\text{Ag}^+ + \text{Sn} \rightarrow \text{Sn}^{2+} + 2\text{Ag}$
4. no; I_2 isn't a strong enough oxidizing agent to spontaneously react with Cl^-

Practice Problems – Determining the Outcome of a Single Replacement Reaction

1. no; I_2 isn't a strong enough oxidizing agent to spontaneously react with F^-
2. yes; $3\text{Cu}^{2+} + 2\text{Al} \rightarrow 2\text{Al}^{3+} + 3\text{Cu}$
3. yes; $\text{Cl}_2 + 2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{Cl}^-$ (* Question revised from complimentary copy)
4. no; Al^{3+} isn't a strong enough oxidizing agent to spontaneously react with Sn

Quick Check

1. Sr
2. Fe^{3+}
3. I^-
4. Sn^{2+}
5. B^-
6. +0.80 V
7. +2.87 V

Practice Problems – Determining the Predominant Redox Reaction

1. $\text{Sn}^{4+} + \text{Ni} \rightarrow \text{Sn}^{2+} + \text{Ni}^{2+}$ (* Question revised from complimentary copy)
2. $3\text{Cu}^{2+} + 2\text{Al} \rightarrow 2\text{Al}^{3+} + 3\text{Cu}$
3. $2\text{Cu}^+ \rightarrow \text{Cu}^{2+} + \text{Cu}$
4. $\text{Cu} + \text{Br}_2 \rightarrow 2\text{Br}^- + \text{Cu}^{2+}$

Practice Problems – Determining a Chemical's Concentration via Redox Titration

- $\text{Cr}_2\text{O}_7^{2-} + 6\text{Fe}^{2+} + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 6\text{Fe}^{3+} + 7\text{H}_2\text{O}$
 - $$0.0150 \text{ L} \times 0.0200 \frac{\text{mol}}{\text{L}} = 3.00 \times 10^{-4} \text{ mol Cr}_2\text{O}_7^{2-}$$

$$3.00 \times 10^{-4} \text{ mol Cr}_2\text{O}_7^{2-} \times \frac{6 \text{ mol Fe}^{2+}}{1 \text{ mol Cr}_2\text{O}_7^{2-}} = 1.80 \times 10^{-3} \text{ mol Fe}^{2+}$$

$$[\text{FeCl}_2] = [\text{Fe}^{2+}] = \frac{1.80 \times 10^{-3} \text{ mol Fe}^{2+}}{0.0200 \text{ L}} = 0.0900 \text{ M}$$
- $$1.00 \text{ g Na}_2\text{C}_2\text{O}_4 \times \frac{1 \text{ mol Na}_2\text{C}_2\text{O}_4}{134.0 \text{ g Na}_2\text{C}_2\text{O}_4} = 7.46 \times 10^{-3} \text{ mol Na}_2\text{C}_2\text{O}_4$$

$$7.46 \times 10^{-3} \text{ mol C}_2\text{O}_4^{2-} \times \frac{2 \text{ mol MnO}_4^-}{5 \text{ mol C}_2\text{O}_4^{2-}} = 2.99 \times 10^{-3} \text{ mol MnO}_4^-$$

$$[\text{KMnO}_4] = [\text{MnO}_4^-] = \frac{2.99 \times 10^{-3} \text{ mol MnO}_4^-}{0.0163 \text{ L}} = 0.183 \text{ M}$$
- $$0.02625 \text{ L} \times 0.500 \frac{\text{mol}}{\text{L}} = 1.3125 \times 10^{-2} \text{ mol Cr}_2\text{O}_7^{2-}$$

$$1.3125 \times 10^{-2} \text{ mol Cr}_2\text{O}_7^{2-} \times \frac{3 \text{ mol CH}_3\text{CH}_2\text{OH}}{2 \text{ mol Cr}_2\text{O}_7^{2-}} = 1.969 \times 10^{-2} \text{ mol CH}_3\text{CH}_2\text{OH}$$

$$[\text{CH}_3\text{CH}_2\text{OH}] = \frac{1.969 \times 10^{-2} \text{ mol CH}_3\text{CH}_2\text{OH}}{0.0100 \text{ L}} = 1.97 \text{ M}$$
 - $$1.97 \frac{\text{mol}}{\text{L}} \times 46.0 \frac{\text{g}}{\text{mol}} \times \frac{1 \text{ mL}}{0.789 \text{ g}} = \frac{115 \text{ mL CH}_3\text{CH}_2\text{OH}}{\text{L beverage}} = 11.5\%$$

Activity: Making an SRP Table

- | Oxidizing Agents | | Reducing Agents |
|-------------------------------|----------------------|-----------------|
| $\text{Q} + \text{e}^-$ | \rightleftharpoons | Q^- |
| $\text{D}^{3+} + \text{e}^-$ | \rightleftharpoons | D^{2+} |
| $\text{P} + \text{e}^-$ | \rightleftharpoons | P^- |
| $\text{M}^{2+} + 2\text{e}^-$ | \rightleftharpoons | M |
| $\text{L}^+ + \text{e}^-$ | \rightleftharpoons | L |
| $\text{C}^{2+} + 2\text{e}^-$ | \rightleftharpoons | C |
- Q^-
- C^{2+}

Review Questions

- no
 - yes
- Either Fe^{3+} , Hg_2^{2+} , Ag^+ , or Hg^{2+}
- yes; $3\text{Mg} + 2\text{Al}^{3+} \rightarrow 3\text{Mg}^{2+} + 2\text{Al}$
 - no
 - yes; $\text{Hg}^{2+} + 2\text{Ag} \rightarrow \text{Hg} + 2\text{Ag}^+$

4.

Metals	Non-metals
bottom right of SRP Table	top left of SRP Table
tend to give electrons	tend to take electrons
give e^- to chemicals above them on the left	take e^- from chemicals below them on the right

- yes; $\text{Fe} + \text{Sn}^{2+} \rightarrow \text{Fe}^{2+} + \text{Sn}$
 - yes; $\text{F}_2 + 2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{F}^-$
 - no
- $2\text{I}^- + 2\text{Fe}^{3+} \rightarrow \text{I}_2 + 2\text{Fe}^{2+}$
 - $\text{Br}_2 + 2\text{Fe}^{2+} \rightarrow 2\text{Br}^- + 2\text{Fe}^{3+}$
- $3\text{Sn}^{2+} + 2\text{NO}_3^- + 8\text{H}^+ \rightarrow 2\text{NO} + 4\text{H}_2\text{O} + 3\text{Sn}^{4+}$
- No. The Fe^{3+} would oxidize and dissolve the Al container.
- reverse
 - forward
- $\text{Mg} + 2\text{H}^+ \rightarrow \text{Mg}^{2+} + \text{H}_2$
- The H^+ in acids is not a strong enough oxidizing agent to oxidize silver but the nitrate ion in nitric acid is.
- $2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{Na}^+ + \text{H}_2 + 2\text{OH}^-$
The OH^- turns the phenolphthalein pink. The H_2 is ignited by the energy released by the exothermic reaction.
- Ag^+

b. Mg

14. Ga^{3+} ; Al is being oxidized by gallium ions therefore Ga^{3+} is a stronger oxidizing agent (higher on the left in the SRP table) than Al^{3+} and so Ga^{3+} has a greater reduction potential than Al^{3+} .

15. a. +1.09V

b. +0.76V

16. a. $\text{Fe} + \text{Co}^{2+} \rightarrow \text{Fe}^{2+} + \text{Co}$

b. $\text{Cu}^+ + \text{Cr}^{2+} \rightarrow \text{Cr}^{3+} + \text{Cu}$ (* Question revised from complimentary copy)

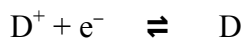
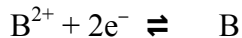
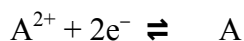
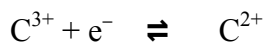
c. $\text{Cu}^{2+} + \text{Sn}^{2+} \rightarrow \text{Sn}^{4+} + \text{Cu}$ (* Question revised from complimentary copy)

17. $\text{Zn} + 2\text{Fe}^{3+} \rightarrow 2\text{Fe}^{2+} + \text{Zn}^{2+}$

18. a. C^{3+}

b. D

SRP Table



19. $0.0328 \text{ L} \times \frac{0.200 \text{ mol}}{\text{L}} = 6.56 \times 10^{-3} \text{ mol Cr}_2\text{O}_7^{2-}$

$$6.56 \times 10^{-3} \text{ mol Cr}_2\text{O}_7^{2-} \times \frac{6 \text{ mol I}^-}{1 \text{ mol Cr}_2\text{O}_7^{2-}} = 3.936 \times 10^{-2} \text{ mol I}^-$$

$$[\text{KI}] = [\text{I}^-] = \frac{3.936 \times 10^{-2} \text{ mol I}^-}{0.0150 \text{ L}} = 2.62 \text{ M}$$

20. $0.105 \text{ g H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}}{126.0 \text{ g H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}} = 8.33 \times 10^{-4} \text{ mol H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$

$$8.33 \times 10^{-4} \text{ mol H}_2\text{C}_2\text{O}_4 \times \frac{2 \text{ mol MnO}_4^-}{5 \text{ mol H}_2\text{C}_2\text{O}_4} = 3.33 \times 10^{-4} \text{ mol MnO}_4^-$$

$$[\text{KMnO}_4] = [\text{MnO}_4^-] = \frac{3.33 \times 10^{-4} \text{ mol MnO}_4^-}{0.0186 \text{ L}} = 0.0179 \text{ M}$$

21. a. $2\text{Cr}_2\text{O}_7^{2-} + \text{C}_2\text{H}_5\text{OH} + 16\text{H}^+ \rightarrow 4\text{Cr}^{3+} + 11\text{H}_2\text{O} + 2\text{CO}_2$

b.

$$0.01015 \text{ L} \times 0.0150 \frac{\text{mol}}{\text{L}} = 1.5225 \times 10^{-4} \text{ mol Cr}_2\text{O}_7^{2-}$$

$$1.5225 \times 10^{-4} \text{ mol Cr}_2\text{O}_7^{2-} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{2 \text{ mol Cr}_2\text{O}_7^{2-}} = 7.6125 \times 10^{-5} \text{ mol C}_2\text{H}_5\text{OH}$$

$$7.6125 \times 10^{-5} \text{ mol C}_2\text{H}_5\text{OH} \times \frac{46.0 \text{ g C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_2\text{H}_5\text{OH}} = 3.50 \times 10^{-3} \text{ g C}_2\text{H}_5\text{OH}$$

$$\frac{3.50 \times 10^{-3} \text{ g C}_2\text{H}_5\text{OH}}{5.00 \text{ g C}_2\text{H}_5\text{OH}} = 0.0700\% \quad \text{Legally impaired now, not before Sep.2010.}$$

$$5.00 \text{ g C}_2\text{H}_5\text{OH}$$

6.4 The Electrochemical Cell

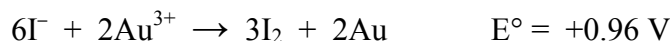
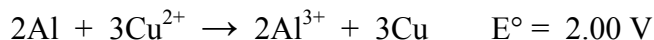
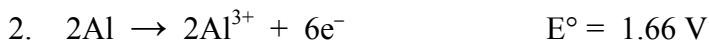
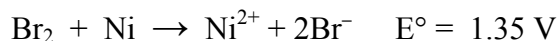
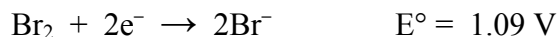
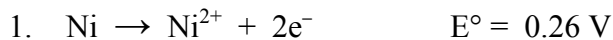
Warm Up

1. ion
2. electron
3. 6.25×10^{18}

Quick Check

1. Cr
2. from right to left or from the Cr electrode to the Ag electrode
3. from right to left or from the Cr/Cr³⁺ half-cell to the Ag/Ag⁺ half-cell
4. Ag

Practice Problems – Determining Standard Cell Potentials (Voltages)



Quick Check

1.

Type	Anode Material	Cathode Material	Electrolyte Medium	Use
Alkaline Cell	Zn powder packed around a brass pin	MnO ₂ (s) and powdered C	a moist paste of KOH	e.g. portable electronics, toys, flashlights
Lead-acid Storage Battery	Pb alloy grids packed with spongy Pb	Pb alloy grids packed with PbO ₂	H ₂ SO ₄ (aq)	automobiles
Fuel Cell	porous carbon	porous carbon	proton exchange membrane	e.g. electric vehicles, space travel,

Quick Check

- hydrated iron (III) oxide or iron(III) oxide monohydrate
- $\frac{1}{2} \text{O}_2(\text{g}) + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{O}$ or $\frac{1}{2} \text{O}_2(\text{g}) + \text{H}_2\text{O} + 2\text{e}^- \rightarrow 2\text{OH}^-$
- impurities or physical stresses or [O₂] or [ions] or surface area
- coating it with paint or grease or cathodic protection or galvanizing it.

Review Questions

1.

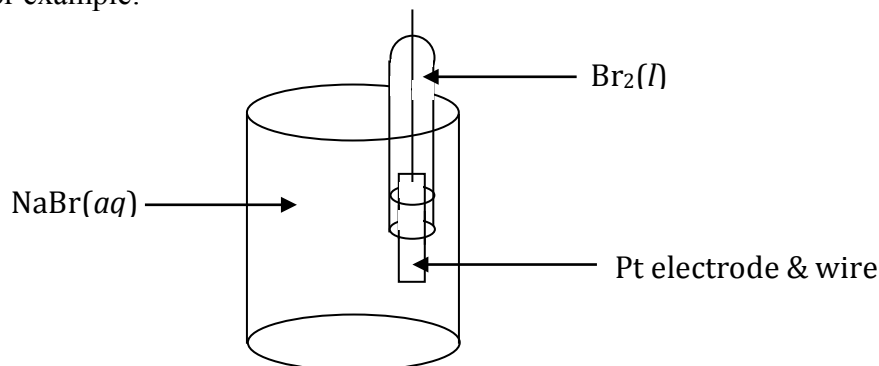
Anode	Cathode
oxidation occurs	reduction occurs
mass decreases	mass increases
attracts anions	attracts cations
electrons flow away	electrons flow towards

- The diagram should show e⁻ flowing through the wire from the Fe/Fe²⁺ half-cell to the Ni/Ni²⁺ half-cell.
 - Fe/Fe²⁺
 - $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$; $\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$
 - Fe labelled anode; Ni labelled cathode

- e. $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^- \quad E^\circ = +0.45 \text{ V}$
 $\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni} \quad E^\circ = -0.26 \text{ V}$

 $\text{Fe} + \text{Ni}^{2+} \rightarrow \text{Fe}^{2+} + \text{Ni} \quad E^\circ = +0.19 \text{ V}$
- f. NH_4^+ flow from the salt bridge into the Ni/Ni^{2+} half-cell
 NO_3^- flow through the salt bridge into the Fe/Fe^{2+} half-cell
- g. The Fe electrode loses mass. The Ni electrode gains mass.

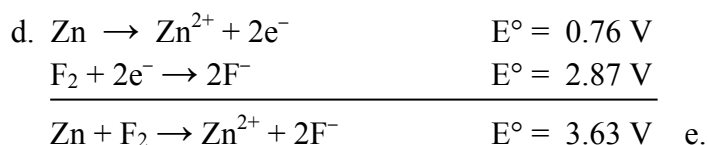
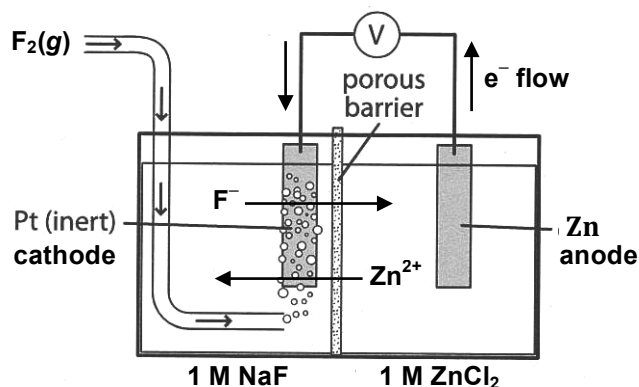
3. for example:



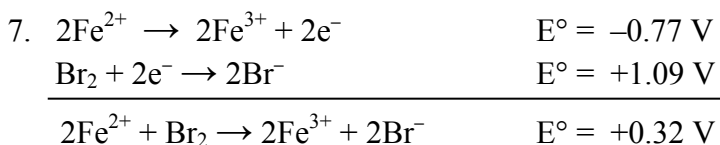
4. a. The diagram should show e^- flowing through the wire from the Cr/Cr^{3+} half-cell to the H_2/H^+ half-cell
- b. H_2/H^+
- c. $\text{Cr} \rightarrow \text{Cr}^{3+} + 3\text{e}^-; 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$
5. a. Cr labelled anode; Pt labelled cathode
- b. $2\text{Cr} \rightarrow 2\text{Cr}^{3+} + 6\text{e}^- \quad E^\circ = 0.74 \text{ V}$
 $6\text{H}^+ + 6\text{e}^- \rightarrow 3\text{H}_2 \quad E^\circ = 0.00 \text{ V}$

 $2\text{Cr} + 6\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 3\text{H}_2 \quad E^\circ = 0.74 \text{ V}$
- c. Cr^{3+} flow through the porous barrier into the H_2/H^+ half-cell
 Cl^- flow through the porous barrier into the Cr/Cr^{3+} half-cell
- d. The Cr electrode loses mass. The Pt electrode's mass is unchanged.

5.

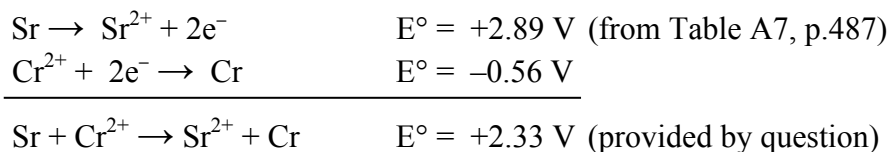


6. it allows ion migration that completes the circuit

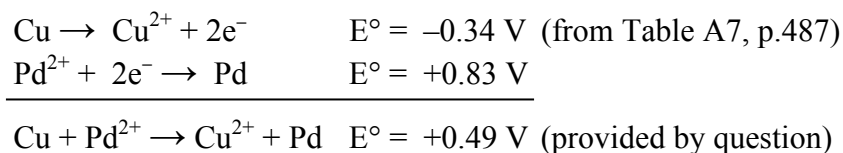


8. a. 2.07 V (the E° of $\text{F}_2 + 2\text{e}^{-} \rightarrow 2\text{F}^{-}$ is 2.07 V greater than the E° of $\text{Ag}^{+} + \text{e}^{-} \rightarrow \text{Ag}$)

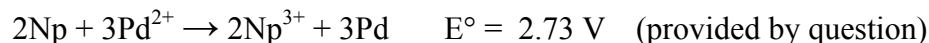
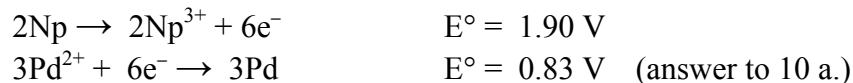
b. -3.17 V (the E° of $\text{Mg}^{2+} + 2\text{e}^{-} \rightarrow \text{Mg}$ is 3.17 V less than the E° of $\text{Ag}^{+} + \text{e}^{-} \rightarrow \text{Ag}$)



10. a. 0.83 V



b. -1.90 V

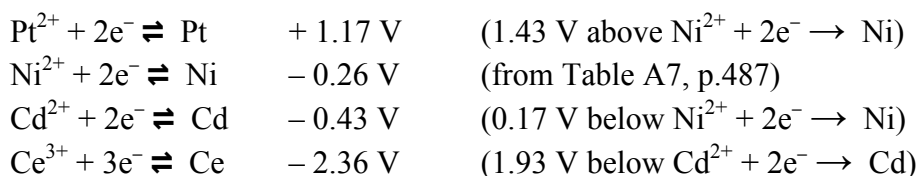


1.90 V is the Standard Oxidation Potential (SOP) of Np therefore the Standard Reduction Potential (SRP) of Np^{3+} is -1.90 V .

11.

		Reduction Half-cell	
		$\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$	$\text{Pt}^{2+} + 2\text{e}^- \rightarrow \text{Pt}$
Oxidation Half-cell	$\text{Pt} \rightarrow \text{Pt}^{2+} + 2\text{e}^-$	-1.60 V	0 V
	$\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^-$	-0.17 V	$+1.43 \text{ V}$
	$\text{Ce} \rightarrow \text{Ce}^{3+} + 3\text{e}^-$	$+1.93 \text{ V}$	$+3.53 \text{ V}$

Construct an SRP Table from the data provided:



12. a. increased voltage (The reactant concentrations are greater than 1 M so the cell's potential will be greater than the standard cell potential.)
 b. 0 V
 c. increased voltage (The product concentrations are less than 1 M because the S^{2-} ions precipitate out some Fe^{2+} so the cell's potential will be greater than the standard cell potential.)
 d. no effect
13. For e.g. Reactants biodegradable. Products are just CO_2 and H_2O . No metal or metal ion reactants or products.
14. Weld a zinc bar to the staircase preferably above the area exposed to ocean spray. The zinc is oxidized more readily than the steel (Fe). The zinc bar transfers electrons through the staircase to the oxidizing agents in the ocean spray thus

leaving the staircase intact.

- If the zinc is oxidized it forms zinc oxide. Zinc oxide forms a hard, impenetrable coating. Should the zinc or the zinc oxide coating be scratched off in some areas, the remaining zinc coating still provides cathodic protection.
- The reactants in electrochemical cells react through electrical contact, via an action-at-a-distance force, without ever coming into physical contact, i.e. colliding.

6.5 The Electrolytic Cell

Warm Up

Device	Energy Conversion	
	From	To
light bulb	electrical energy	light energy
photocell	light energy	electrical energy

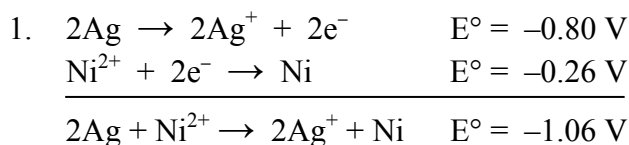
speaker	electrical energy	sound energy
microphone	sound energy	electrical energy

heating element	electrical energy	heat energy
thermocouple	heat energy	electrical energy

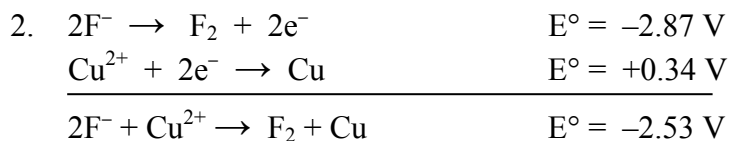
motor	electrical energy	mechanical energy
generator	mechanical energy	electrical energy

electrolytic cell	electrical energy	chemical energy
electrochemical cell	chemical energy	electrical energy

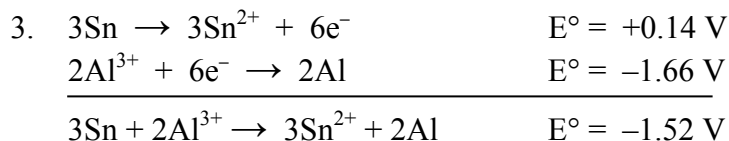
Practice Problems – Predicting the Voltage Required to Operate an Electrolytic Cell



A voltage of at least 1.06 V would be required to operate this cell.



A voltage of at least 2.53 V would be required to operate this cell.



A voltage of at least 1.52 V would be required to operate this cell.

Quick Check

1. electrochemical cell
2. electrolytic cell
3. both
4. electrochemical cell
5. both
6. electrolytic cell
7. electrochemical cell

Practice Problems – Predicting the Half-Reactions That Will Occur in an Electrolytic Cell

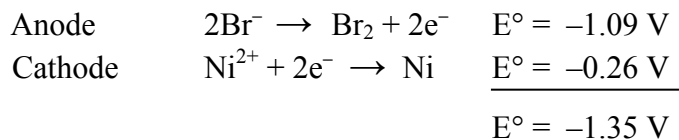
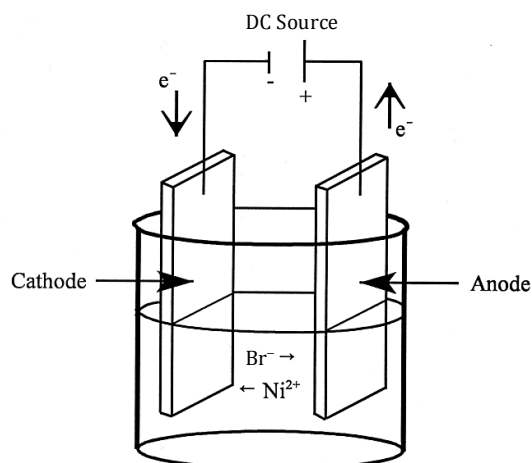
1. a. Type 1

$$\begin{array}{l}
 2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^- \\
 \text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}
 \end{array}$$
- b. Type 2

$$\begin{array}{l}
 \text{H}_2\text{O} \rightarrow \frac{1}{2} \text{O}_2 + 2\text{H}^+ + 2\text{e}^- \\
 2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2 \text{OH}^-
 \end{array}$$
- c. Type 3

$$\begin{array}{l}
 \text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^- \\
 2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2 \text{OH}^-
 \end{array}$$

2.



A voltage of at least 1.35 V would be required to operate this cell.

Quick Check

1. the electrolytic recovery of a metal from a solution containing its ions.
2. aluminum oxide dissolves in molten cryolite ($\text{Na}_3\text{AlF}_6(l)$)
3. sodium hydroxide and chlorine
4. impressed current cathodic protection

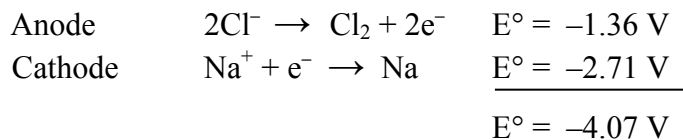
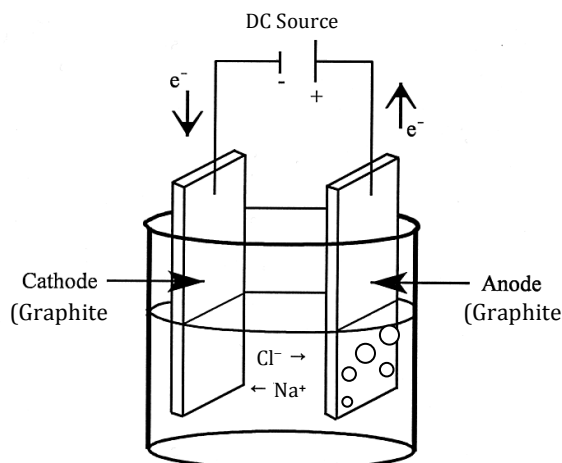
Activity: Location, Location, Location

- | | |
|--|---|
| <ol style="list-style-type: none"> 1. For example: proximity to source of ore
proximity to shipping port
proximity to market
inexpensive electricity
available work force
inexpensive labour
stable political climate
democracy | <ol style="list-style-type: none"> 3. <input checked="" type="checkbox"/> Brazil
<input checked="" type="checkbox"/> deep water port
<input checked="" type="checkbox"/> western USA, Japan
<input checked="" type="checkbox"/> B.C. Hydro
<input checked="" type="checkbox"/> company town
<input checked="" type="checkbox"/> strong labour unions
<input checked="" type="checkbox"/> Canada – stable |
|--|---|

Review Questions

1. citric acid (weak electrolyte), sodium chloride, sodium citrate, and potassium dihydrogen phosphate

2.

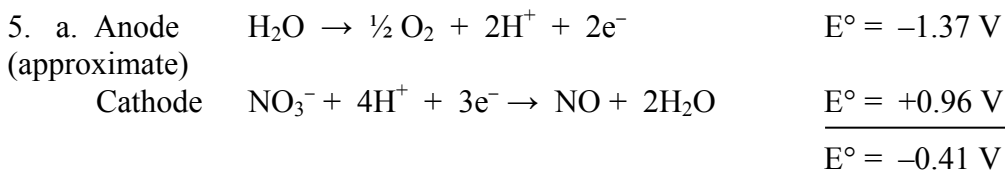


A voltage of at least 4.07 V would be required to operate this cell.

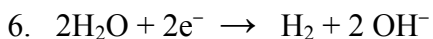
3. because the products (Na and Cl₂) spontaneously react with each other

4.

	Cell Type (1,2,3)	Electrolyte	Anode/Cathode	Products	
				Anode	Cathode
a.	1	NaCl(l)	Pt / Pt	Cl ₂	Na
b.	2	NaCl(aq)	Pt / C	Cl ₂	H ₂ + 2OH ⁻
c.	2	CuBr ₂ (aq)	C / C	Br ₂	Cu
d.	2	AlF ₃ (aq)	C / C	½O ₂ + 2H ⁺	H ₂ + 2OH ⁻
e.	3	CuCl ₂ (aq)	Cu / Cu	Cu ²⁺	Cu

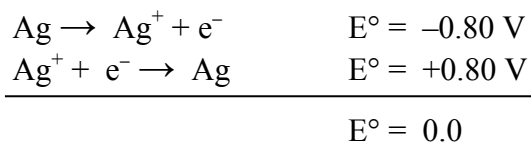
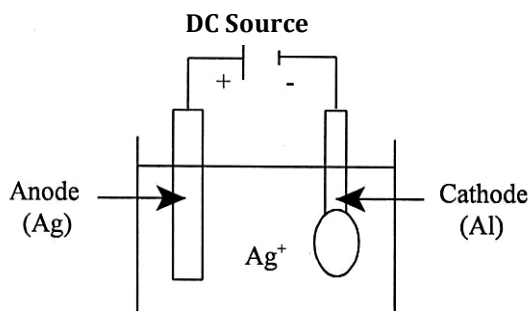


b. A voltage of at least 0.41 V would be required to operate this cell.



7. In electrorefining, the metallic ions reduced at the cathode are replaced by oxidizing an impure anode.
8. Uneconomical to electrolyze $\text{Al}_2\text{O}_3(l)$ because aluminum oxide has a high melting point. Impossible to produce Al by electrolyzing $\text{Al}_2\text{O}_3(aq)$ because water is a stronger oxidizing agent than Al^{3+} .
9. In impressed current cathodic protection the metal being protected acts as the cathode in an electrolytic cell (non-spontaneous redox reaction) whereas in galvanic cathodic protection the metal being protected acts as the cathode in an electrochemical cell (spontaneous redox reaction).

10.



In theory, it should require no voltage to operate this cell because no net reaction occurs. In practice, a small voltage is required to overcome the internal resistance of the cell.

11. a. Type 1 Cell. Only one chemical, water, can be oxidized or reduced. Water is molten H_2O .
- b. Any salt containing ions which are weaker reducing agents and weaker oxidizing agents than water; for example, sodium sulphate.

12. 0.75 V

