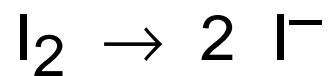


6.3-110

In the half reaction



the iodine is

- (a) reduced
- (b) oxidized
- (c) neither of the above

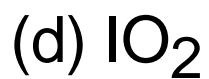
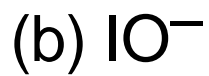
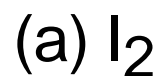
6.3-120

Vitamin C is an "antioxidant". This is because it

- (a) oxidizes readily
- (b) is an oxidizing agent
- (c) neither of the above

6.3-130

In which of the following is the oxidation number of iodine greatest?



6.3-140

If a reducing agent is to be used to generate the various oxidation states of vanadium, one should start with

- (a) VO_3^- (yellow)
- (b) VO^{2+} (green)
- (c) V^{3+} (blue)
- (d) V^{2+} (violet)

6.3-160

In the reaction

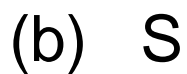


oxygen is

- (a) oxidized
- (b) reduced
- (c) neither oxidized or reduced
- (d) both oxidized and reduced

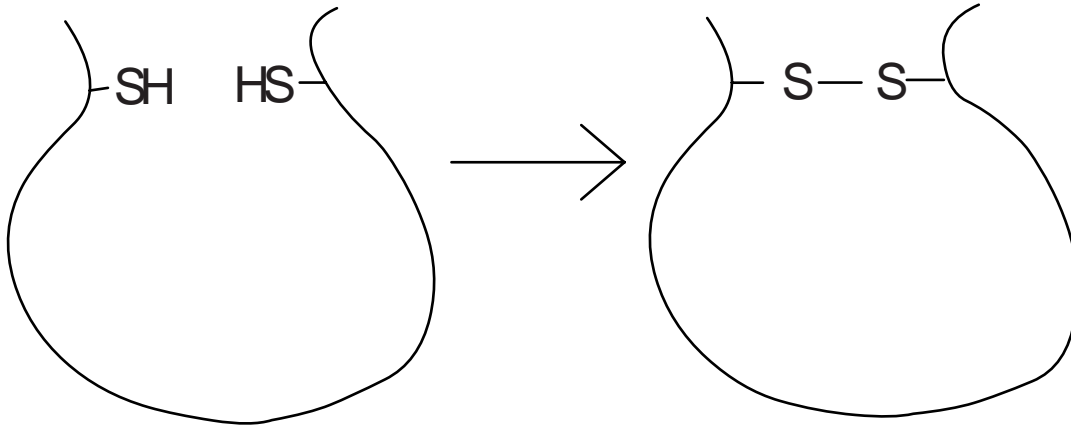
6.3-170

When two of the following compounds are mixed, a redox reaction occurs in which the third is formed. Which is the product?



6.3-180

Protein folding is often stabilized by the formation of disulfide bonds between cysteine side chains. In this reaction



the sulfur is

- (a) reduced
- (b) oxidized
- (c) neither of the above

6.3-210

In balancing the oxidation of methanol to formic acid



the number of electrons that must be added to the right side is

- | | | | |
|-----|---|-----|---|
| (a) | 0 | (b) | 2 |
| (c) | 4 | (d) | 6 |

6.3-220

In balancing the half-reaction

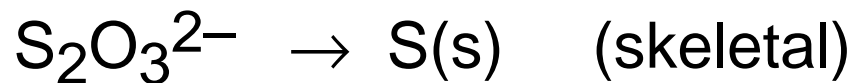


the number of electrons that must be added is

- (a) 0
- (b) 2 on the left
- (c) 2 on the right

6.3-230

In balancing the half reaction

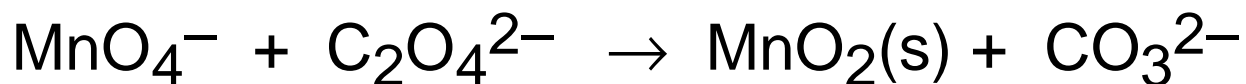


the number of electrons that must be added is

- (a) 2 on the left
- (b) 2 on the right
- (c) 4 on the left
- (d) 4 on the right

6.3-260

For the reaction between permanganate ion and oxalate ion in basic solution, the unbalanced equation is

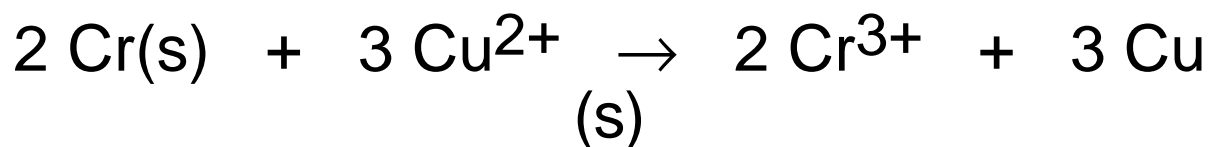


When this equation is balanced, the number of OH^- ions is

- (a) 2 on the left
- (b) 2 on the right
- (c) 4 on the left
- (d) 4 on the right

6.3-310

The line representation for a voltaic cell utilizing the reaction

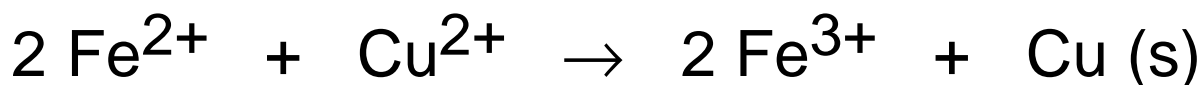


is

- (a) $\text{Cr(s)} \mid \text{Cu}^{2+} \parallel \text{Cr}^{3+} \mid \text{Cu(s)}$
- (b) $\text{Cr(s)} \mid \text{Cr}^{3+} \parallel \text{Cu}^{2+} \mid \text{Cu(s)}$
- (c) $\text{Cu(s)} \mid \text{Cu}^{2+} \parallel \text{Cr}^{3+} \mid \text{Cr(s)}$
- (d) $\text{Cu(s)} \mid \text{Cr}^{3+} \parallel \text{Cu}^{2+} \mid \text{Cr(s)}$

6.3-320

The line representation for a voltaic cell utilizing the reaction



is

- (a) $\text{Fe}^{2+} | \text{Fe}^{3+} || \text{Cu}^{2+} | \text{Cu(s)}$
- (b) $\text{Fe}^{3+} | \text{Fe}^{2+} || \text{Cu}^{2+} | \text{Cu(s)}$
- (c) $\text{Cu(s)} | \text{Fe}^{3+}, \text{Cu}^{2+} || \text{Fe}^{2+} | \text{Cu(s)}$
- (d) $\text{Pt(s)} | \text{Fe}^{3+}, \text{Fe}^{2+} || \text{Cu}^{2+} | \text{Cu(s)}$

6.3-360

Consider the voltaic cell



The anode reaction produces

- | | |
|----------------------|----------------------|
| (a) Sn^{2+} | (b) Fe^{2+} |
| (c) Sn^{4+} | (d) Fe^{3+} |

6.3-370

Consider the voltaic cell



The cathode reaction produces



6.3-390

When a voltaic cell is running, cations in the salt bridge

- (a) move into the cathode compartment
- (b) move into the anode compartment
- (c) stay put

6.3-410

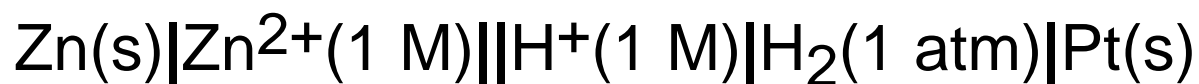
$\epsilon^{\circ}_{H^+|H_2}$ is

- (a) positive
- (b) negative
- (c) zero

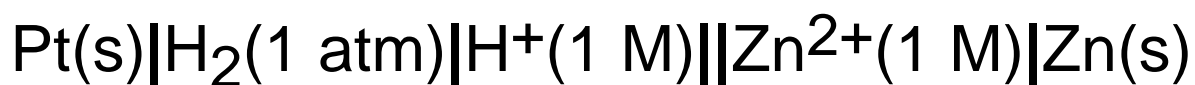
6.3-420

$\varepsilon_{\text{Zn}^{2+}|\text{Zn}}^0$ is the EMF of the cell

(a)



(b)



6.3-430

For the cell



the EMF is

- (a) $+ \varepsilon_{\text{Zn}^{2+}|\text{Zn}}^{\circ} - \varepsilon_{\text{Cl}_2|\text{Cl}^{-}}^{\circ}$
- (b) $+ \varepsilon_{\text{Zn}^{2+}|\text{Zn}}^{\circ} + \varepsilon_{\text{Cl}_2|\text{Cl}^{-}}^{\circ}$
- (c) $- \varepsilon_{\text{Zn}^{2+}|\text{Zn}}^{\circ} - \varepsilon_{\text{Cl}_2|\text{Cl}^{-}}^{\circ}$
- (d) $- \varepsilon_{\text{Zn}^{2+}|\text{Zn}}^{\circ} + \varepsilon_{\text{Cl}_2|\text{Cl}^{-}}^{\circ}$

6.3-450

If the cells



and



are connected in series, the total EMF is

$$(a) \quad 2 \varepsilon_{\text{Pb}^{2+}|\text{Pb}}^{\circ} - \varepsilon_{\text{Zn}^{2+}|\text{Zn}}^{\circ} - \varepsilon_{\text{Cl}_2|\text{Cl}^{-}}^{\circ}$$

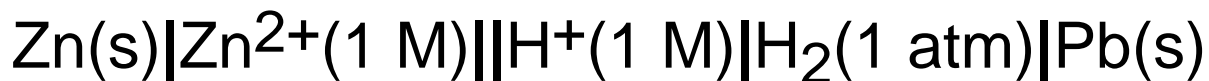
$$(b) \quad + \varepsilon_{\text{Zn}^{2+}|\text{Zn}}^{\circ} + \varepsilon_{\text{Cl}_2|\text{Cl}^{-}}^{\circ} - 2 \varepsilon_{\text{Pb}^{2+}|\text{Pb}}^{\circ}$$

$$(c) \quad + \varepsilon_{\text{Zn}^{2+}|\text{Zn}}^{\circ} - \varepsilon_{\text{Cl}_2|\text{Cl}^{-}}^{\circ}$$

$$(d) \quad - \varepsilon_{\text{Zn}^{2+}|\text{Zn}}^{\circ} + \varepsilon_{\text{Cl}_2|\text{Cl}^{-}}^{\circ}$$

6.3-460

If the cell



has an EMF of 0.763 V and the cell



has an EMF of 0.427 V, then $\varepsilon_{\text{Ti}^{+}|\text{Ti}}^{\circ} =$

- (a) (+ 0.763 + 0.427) V
- (b) (- 0.763 + 0.427) V
- (c) (+ 0.763 - 0.427) V
- (d) (- 0.763 - 0.427) V

6.3-510

Given that

$$\varepsilon_{Ni^{2+}|Ni}^{\circ} = -0.23 \text{ V}$$

and

$$\varepsilon_{Zn^{2+}|Zn}^{\circ} = -0.763 \text{ V}$$

it is expected that

- (a) Ni will reduce Zn^{2+} to Zn
- (b) Zn will reduce Ni^{2+} to Ni
- (c) neither of the above

6.3-530

Given that

$$\varepsilon_{\text{Cu}^{2+}|\text{Cu}}^{\circ} = + 0.340 \text{ V}$$

and

$$\varepsilon_{\text{Zn}^{2+}|\text{Zn}}^{\circ} = - 0.763 \text{ V}$$

the cell



will run

(a) forward

(b) backward

6.3-550

From the standard electrode potentials

$$\mathcal{E}_{Fe^{3+}|Fe^{2+}}^{\circ} = + 0.770 \text{ V}$$

and

$$\mathcal{E}_{Fe^{3+}|Fe}^{\circ} = - 0.036 \text{ V}$$

and

$$\mathcal{E}_{Fe^{2+}|Fe}^{\circ} = - 0.44 \text{ V}$$

it follows that the disproportionation of Fe^{2+}

(a) is (b) is not
spontaneous under standard conditions.

6.3-560

From the standard electrode potentials

$$\varepsilon_{\text{Cu}^{2+}|\text{Cu}^+}^{\circ} = + 0.153 \text{ V}$$

and

$$\varepsilon_{\text{Cu}^{2+}|\text{Cu}}^{\circ} = + 0.340 \text{ V}$$

and

$$\varepsilon_{\text{Cu}^+|\text{Cu}}^{\circ} = + 0.521 \text{ V}$$

it follows that the disproportionation of Cu^+

(a) is (b) is not
spontaneous under standard conditions.

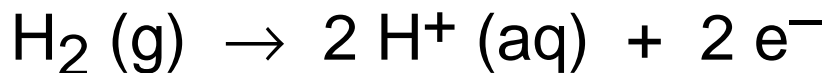
6.3-610

Which of the following statements about a galvanic cell is true?

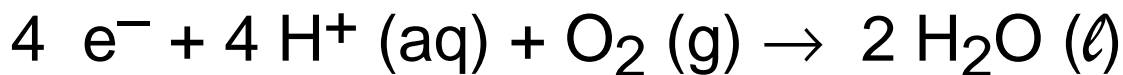
- (a) If $\Delta\mathcal{E} < 0$, then $\Delta G > 0$
- (b) If $\Delta\mathcal{E}^\circ > 0$, then $\Delta G > 0$
- (c) If $\Delta\mathcal{E}^\circ < 0$, then $K_{\text{eq}} > 1$
- (d) If $\Delta\mathcal{E} > 0$, then $K_{\text{eq}} > 1$

6.3-620

A fuel cell involves the anode reaction



and the cathode reaction



Since $\Delta G_f^\circ(\text{H}_2\text{O} (\ell)) = -237.18 \text{ kJ/mol}$,

(a) $\Delta \mathcal{E}^\circ = + (237.18 \text{ kJ/mol}) / 2$ ~~✓~~

(b) $\Delta \mathcal{E}^\circ = + (237.18 \text{ kJ/mol}) / 4$ ~~✓~~

(c) $\Delta \mathcal{E}^\circ = - (237.18 \text{ kJ/mol}) / 2$ ~~✓~~

(d) $\Delta \mathcal{E}^\circ = - (237.18 \text{ kJ/mol}) / 4$ ~~✓~~

6.3-630

Consider the cell



If y is decreased, the EMF of the cell will

- (a) increase
- (b) decrease
- (c) not change

6.3-650

When a strip of Cu and a strip of Zn are poked into a piece of fruit or vegetable (or suspended in a beverage) without direct contact, a wire connecting them is found to carry current. Given that

$$\varepsilon_{\text{Zn}^{2+}|\text{Zn}}^{\circ} = -0.763 \text{ V} \quad \& \quad \varepsilon_{\text{Cu}^{2+}|\text{Cu}}^{\circ} = +0.340 \text{ V}$$

the correct representation of the cell is

- (a) $\text{Cu(s)} \mid \text{Cu}^{2+}(\text{aq}) \parallel \text{Zn}^{2+}(\text{aq}) \mid \text{Zn(s)}$
- (b) $\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}) \parallel \text{Cu}^{2+}(\text{aq}) \mid \text{Cu(s)}$
- (c) $\text{Cu(s)} \mid \text{H}_2(\text{g}) \mid \text{H}^+(\text{aq}), \text{Zn}^{2+}(\text{aq}) \mid \text{Zn(s)}$
- (d) $\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}), \text{H}^+(\text{aq}) \mid \text{H}_2(\text{g}) \mid \text{Cu(s)}$

A more acidic fruit, vegetable or beverage will

- (a) enhance
 - (b) reduce
 - (c) have no effect on
- the cell potential

As the $[\text{Zn}^{2+}]$ in the cell increases, the cell potential will

- (a) increase
- (b) decrease
- (c) remain the same

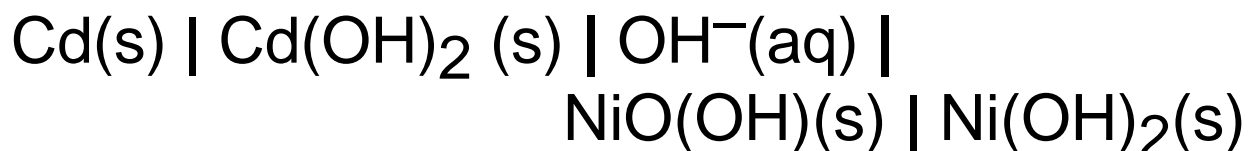
If the size of the Zn electrode is increased, the cell potential will

- (a) increase
- (b) decrease
- (c) remain the same

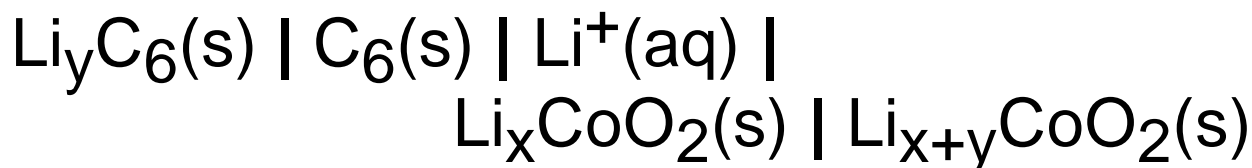
6.3-660

Which of the following commercial cells does not have a constant discharge voltage?

(a)



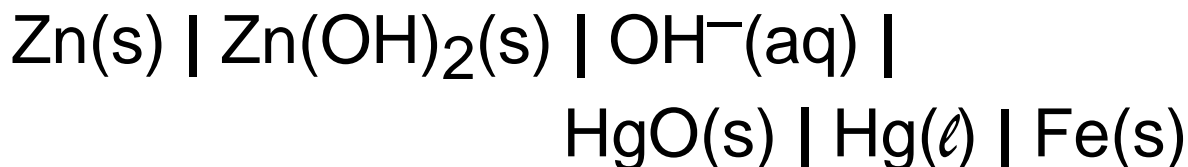
(b)



(c)

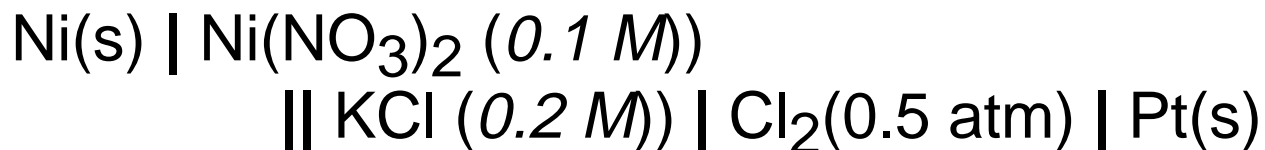


(d)



6.3-670

For the cell



the value of the reaction quotient, Q, is

(a) $(0.5) (0.1) / (0.2)^2$

(b) $(0.1) (0.2) / (0.5)^2$

(c) $(0.5) (0.1)^2 / (0.2)$

(d) $(0.1) (0.2)^2 / (0.5)$

and

(a) $\Delta\mathcal{E} = \Delta\mathcal{E}^\circ + (RT/2\cancel{F}) \ln Q$

(b) $\Delta\mathcal{E} = \Delta\mathcal{E}^\circ - (RT/2\cancel{F}) \ln Q$

(c) $\Delta\mathcal{E} = \Delta\mathcal{E}^\circ + (RT/\cancel{F}) \ln Q$

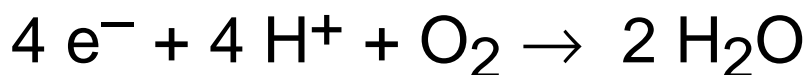
(d) $\Delta\mathcal{E} = \Delta\mathcal{E}^\circ - (RT/\cancel{F}) \ln Q$

6.3-690

A methanol fuel cell involves the anode reaction



and the cathode reaction



the voltage of this cell is

$$(a) \quad \Delta \mathcal{E}^\circ - (RT/6\cancel{F}) \ln \left\{ \frac{p_{\text{CO}_2}}{p_{\text{O}_2} [\text{CH}_3\text{OH}]} \right\}$$

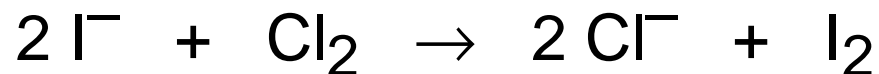
$$(b) \quad \Delta \mathcal{E}^\circ - (RT/4\cancel{F}) \ln \left\{ \frac{p_{\text{CO}_2}}{p_{\text{O}_2} [\text{CH}_3\text{OH}]} \right\}$$

$$(c) \quad \Delta \mathcal{E}^\circ - (RT/12\cancel{F}) \ln \left\{ \frac{p_{\text{CO}_2}^2}{p_{\text{O}_2}^3 [\text{CH}_3\text{OH}]^2} \right\}$$

$$(d) \quad \Delta \mathcal{E}^\circ - (RT/12\cancel{F}) \ln \left\{ \frac{p_{\text{CO}_2}^3}{p_{\text{O}_2}^2 [\text{CH}_3\text{OH}]^3} \right\}$$

6.3-710

For the reaction



the standard cell voltage is 0.82 V.

The equilibrium constant for this reaction is given by $RT \ln K_{\text{eq}} =$

(a) $1 \mathcal{F} (0.82 \text{ V})$

(b) $-1 \mathcal{F} (0.82 \text{ V})$

(c) $2 \mathcal{F} (0.82 \text{ V})$

(d) $-2 \mathcal{F} (0.82 \text{ V})$

6.3-730

Given that

$$\varepsilon_{AgCl|Ag}^0 = + 0.2223 \text{ V}$$

$$\varepsilon_{AgBr|Ag}^0 = + 0.0713 \text{ V}$$

$$\varepsilon_{AgI|Ag}^0 = - 0.152 \text{ V}$$

$$\varepsilon_{Ag^+|Ag}^0 = + 0.799 \text{ V}$$

which has the smallest K_{sp} ?

- (a) AgCl (b) AgBr (c) AgI

$K_{sp}(AgBr) =$

(a) $\exp [\cancel{F} (-0.799+0.0713) \text{ V} / RT]$

(b) $\exp [\cancel{F} (-0.799-0.0713) \text{ V} / RT]$

(c) $\exp [RT / \cancel{F} (-0.799+0.0713) \text{ V}]$

(d) $\exp [RT / \cancel{F} (-0.799-0.0713) \text{ V}]$

6.3-810

Thomas Edison's first electricity meter was based on changes in mass of two parallel Zn plates immersed in zinc sulfate solution. In this device, transfer of a mole of Zn from one plate to the other signifies passage of

(a) 1 C

(b) 1 ~~F~~

(c) 2 C

(d) 2 ~~F~~

6.3-840

Given that the charge of a mole of electrons is $96,485 \text{ C} = 1 \text{ F}$, plating 1 mole of chromium from chromic acid (H_2CrO_4) using a 200 A current will take

(a) $3 (96,485) / 200$

(b) $200 / 3 (96,485)$

(c) $6 (96,485) / 200$

(d) $200 / 6 (96,485)$

seconds.

6.3-860

To plate Cr(s) from aqueous solution under standard conditions, one can start with

- (a) Cr³⁺
- (b) Cr²⁺
- (c) either
- (d) neither

given that $\varepsilon_{Cr^{3+}|Cr}^{\circ} = -0.74 \text{ V}$

$$\varepsilon_{Cr^{2+}|Cr}^{\circ} = -0.91 \text{ V}$$

$$\varepsilon_{H_2O|H_2}^{\circ} = -0.8277 \text{ V}$$

6.3-920

When a dilute solution of Li_2SO_4 is electrolyzed, the products formed at the anode and cathode respectively are

- (a) S and Li
- (b) O_2 and Li
- (c) SO_2 and H_2
- (d) O_2 and H_2

Possibly useful information:

$$\varepsilon_{\text{O}_2|\text{H}_2\text{O}}^{\circ} = + 1.229 \text{ V}$$

$$\varepsilon_{\text{H}_2\text{O}|\text{H}_2}^{\circ} = - 0.8277 \text{ V}$$

$$\varepsilon_{\text{Li}^+|\text{Li}}^{\circ} = - 3.405 \text{ V}$$

6.3-940

Electricity from a fuel cell that consumes H_2 and O_2 is used to drive the electrolysis of $\text{H}_2\text{C}_2\text{O}_4$ to CO_2 and H_2 . Assuming that all the gases are held under the same conditions, the volume of O_2 consumed is

- (a) one fourth
- (b) half
- (c) the same as
- (d) twice

the volume of CO_2 produced.