001 ________________ ________________ ________________________
version last name first name signature

Practice Exam 2

TTh Classes · Spring 2016

REMEMBER: Bubble in ALL Bubblesheet information!

This includes your first and last name, your UTEID, and your version number.

Please refer to the back of the bubble sheet for more info.

$F = 96{,}485$ C/mol of e^{-1}

1 Ampere = 1 C/s

$N_a = 6.022 \times 10^{23}$

$E^\circ_{cell} = E^\circ_{cat} - E^\circ_{an}$

$E_{cell} = E^\circ_{cell} - \left(\frac{0.05916}{\text{n}}\right) \cdot \log Q$

$\Delta G^\circ = -n \cdot F \cdot E^\circ_{cell}$

Standand Potentials at 25° C	E°
$F_2(g) + 2\,e^- \rightleftharpoons 2\,F^-(aq)$	+2.87 V
$Au^{3+}(aq) + 3\,e^- \rightleftharpoons Au(s)$	+1.40 V
$Cl_2(g) + 2\,e^- \rightleftharpoons 2\,Cl^-(aq)$	+1.36 V
$Pt^{2+}(aq) + 2\,e^- \rightleftharpoons Pt(s)$	+1.20 V
$Br_2(g) + 2\,e^- \rightleftharpoons 2\,Br^-(aq)$	+1.08 V
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80 V
$Cu^{2+}(aq) + 2\,e^- \rightleftharpoons Cu(s)$	+0.34 V
$2\,H^+(aq) + 2\,e^- \rightleftharpoons H_2(g)$	**0.000** V
$Fe^{3+}(aq) + 3\,e^- \rightleftharpoons Fe(s)$	−0.040 V
$Ni^{2+}(aq) + 2\,e^- \rightleftharpoons Ni(s)$	−0.236 V
$Cd^{2+}(aq) + 2\,e^- \rightleftharpoons Cd(s)$	−0.40 V
$Fe^{2+}(aq) + 2\,e^- \rightleftharpoons Fe(s)$	−0.44 V
$Cr^{3+}(aq) + 3\,e^- \rightleftharpoons Cr(s)$	−0.74 V
$Cr^{2+}(aq) + 2\,e^- \rightleftharpoons Cr(s)$	−0.91 V
$Al^{3+}(aq) + 3\,e^- \rightleftharpoons Al(s)$	−1.66 V
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	−3.040 V

This print-out should have 25 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

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001 5.0 points

In the reaction

$$2\,MnO_4^- + 16\,H^+ + 10\,I^- \rightarrow 2\,Mn^{2+} + 8\,H_2O + 5\,I_2$$

of permanganate anion with iodide in acidic solution, what has been reduced and what is the change in oxidation number?

1. hydrogen, from +1 to 0

2. manganese, from +7 to +2 **correct**

3. iodine, from −1 to 0

4. iodine, from −10 to −5

5. oxygen, from −1 to −2

Explanation:

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002 5.0 points

Silver is plated on copper by immersing a piece of copper into a solution containing silver(I) ions. In the plating reaction, copper

1. is reduced and is the reducing agent.

2. is reduced and is the oxidizing agent.

3. is oxidized and is the reducing agent. **correct**

4. is oxidized and is the oxidizing agent.

Explanation:

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003 5.0 points

If the standard potentials for the couples $Fe^{3+} \,|\, Fe^{2+}$, $MnO_4^{2-}, H^+ \,|\, Mn^{2+}, H_2O$, $Zn^{2+} \,|\, Zn$, $V^{3+} \,|\, V^{2+}$, and $Br_2 \,|\, Br^-$ are +0.77, +1.51, − 0.76, − 0.26, and +1.09 V, respectively, which is the strongest oxidizing agent?

1. Mn^{2+}

2. Zn^{2+}

3. MnO_4^{2-} **correct**

4. Fe^{3+}

5. Br_2

Explanation:

Each value shown is a standard reduction potential value. It measures the relative ease by which the given reduction process happens. The more positive the number the more likely the process. We are asked which is the strongest OXIDIZING AGENT. An oxidizing agent will itself be reduced. The species which is most likely to be reduced of all listed is MnO_4^{2-} since it is the species with the biggest positive number. Note MnO_4^{2-} has Mn with an oxidation state of 6+ to balance four -2 for the O's and the overall charge, and it is being reduced to Mn^{2+}.

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004 5.0 points

Using the set of smallest whole number coefficients to balance the redox equation

$$MnO_4^- + NO_2^- \rightarrow MnO_2 + NO_3^-$$

in basic solution, you get

1. $1\,H_2O$ on the right.

2. $3\,H_2O$ on the left.

3. $3\,OH^-$ on the left.

4. $2\,OH^-$ on the right. **correct**

Explanation:

The oxidation number of N changes from +3 to +5, so N is oxidized. The oxidation number of Mn changes from +7 to +4, so Mn is reduced. We set up oxidation and reduction half-reactions:

Red: $MnO_4^- \rightarrow MnO_2$

Oxid: $NO_2^- \rightarrow NO_3^-$

Mn and N atoms are balanced. Since this is a basic solution, we use H_2O and OH^- to balance O and H atoms, adding the OH^- to the side needing oxygen:

Red: $2\,H_2O + MnO_4^- \to MnO_2 + 4\,OH^-$
Oxid: $2\,OH^- + NO_2^- \to NO_3^- + H_2O$

We balance the total charge in each half-reaction by adding electrons. In the preceding reduction reaction there is a total charge of -1 on the left and -4 on the right. Three electrons are added to the left:

Red: $3\,e^- + 2\,H_2O + MnO_4^- \to MnO_2 + 4\,OH^-$
Oxid: $2\,OH^- + NO_2^- \to NO_3^- + H_2O + 2\,e^-$

The number of electrons gained by Mn must equal the number of electrons lost by N. We multiply the reduction reaction by 2 and the oxidation reaction by 3 to balance the electrons:

Red: $6\,e^- + 4\,H_2O + 2\,MnO_4^- \to 2\,MnO_2 + 8\,OH^-$
Oxid: $6\,OH^- + 3\,NO_2^- \to 3\,NO_3^- + 3\,H_2O + 6\,e^-$

Adding the half-reactions gives

$$4\,H_2O + 2\,MnO_4^- + 6\,OH^- + 3\,NO_2^- \to 2\,MnO_2 + 8\,OH^- + 3\,NO_3^- + 3\,H_2O$$

Canceling like terms given the overall balanced equation

$$H_2O + 2\,MnO_4^- + 3\,NO_2^- \to 2\,MnO_2 + 2\,OH^- + 3\,NO_3^-$$

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005 5.0 points

Balance the net ionic equation

$$MnO_4^- + I^- \to I_2 + MnO_2$$

in basic solution. What is the sum of the coefficients for ONLY THE REACTANT SIDE of the equation?

1. 14

2. 10

3. 8

4. 12 **correct**

5. 16

6. 21

7. 9

Explanation:

The oxidation number of I changes from -1 to 0, so I is oxidized. The oxidation number of Mn changes from $+7$ to $+4$, so Mn is reduced. We set up oxidation and reduction half reactions:

Oxidation: $I^- \to I_2$
Reduction: $MnO_4^- \to MnO_2$

Mn atoms are balanced. We need two I^- ions to balance I:

Oxidation: $2\,I^- \to I_2$

In basic solution we use H_2O and OH^- to balance O and H atoms, adding the OH^- to the side needing oxygen:

Red: $2\,H_2O + MnO_4^- \to MnO_2 + 4\,OH^-$

Next we balance the total charge by adding electrons. In the reduction reaction thus far there is a total charge of -1 on the left and -4 on the right. Three electrons are added to the left:

Oxid: $2\,I^- \to I_2 + 2\,e^-$
Red: $3\,e^- + 2\,H_2O + MnO_4^- \to MnO_2 + 4\,OH^-$

The number of electrons gained by Mn must equal to the number of electrons lost by I. We multiply the oxidation reaction by 3 and the reduction reaction by 2 to balance the electrons:

Oxid: $6\,I^- \to 3\,I_2 + 6\,e^-$
Red: $6\,e^- + 4\,H_2O + 2\,MnO_4^- \to 2\,MnO_2 + 8\,OH^-$

Adding the half-reactions give the balanced equation

$$6\ I^- + 4\ H_2O + 2\ MnO_4^- \to 3\ I_2 + 2\ MnO_2 + 8\ OH^-$$

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006 5.0 points

Far a battery, the cathode is the (positive/negative) terminal and the electrons flow through the external circuit from (anode to cathode/cathode to anode).

1. negative, anode to cathode

2. positive, cathode to anode

3. negative, cathode to anode

4. positive, anode to cathode **correct**

Explanation:

For a discharging battery, the cathode is the positive terminal. The cathode *always* has an inward flow of electrons, because it is by definition the site of reduction.

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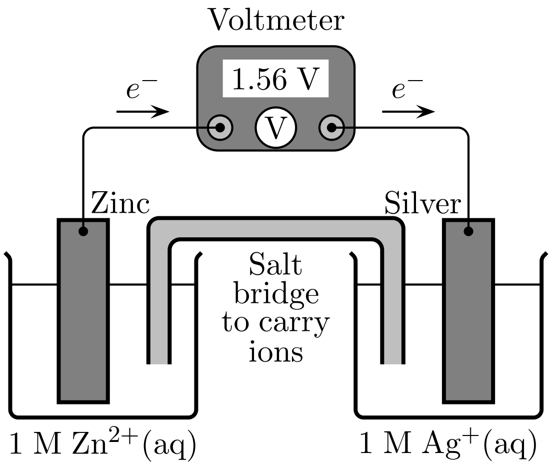
007 5.0 points



In this electrochemical cell, what is the cathode?

1. the Ag^+(aq) ions in the 1 M solution

2. the Zn^{2+}(aq) ions in the 1 M solution

3. the solid silver electrode **correct**

4. the solid zinc electrode

Explanation:

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

Reduction occurs at the cathode. In this cell the reduction half reaction is

$$Ag^+(aq) + e^- \rightarrow Ag(s)\,.$$

Ag^+ cations are attracted to the solid Ag electrode where they are reduced to Ag(s).

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008 5.0 points

What is the primary difference between a voltaic and an electrolytic cell?

1. The reaction in a voltaic cell is spontaneous; the one in an electrolytic cell is nonspontaneous. **correct**

2. Voltaic cells are used only in laboratories; electrolytic cells are used only in industrial applications.

3. The voltaic cell is a real reaction that is used in everyday applications; the electrolytic cell is a theoretical reaction that does not have any real-life applications.

4. The electrolytic cell can be used as a power supply; the voltaic cell must itself have power supplied in order to function.

5. The reaction in a voltaic cell is exothermic; the one in an electrolytic cell is endothermic.

Explanation:

Both types of cells are used in many applications. A rechargable battery switches between being a voltaic cell (when being used) to an electrochemical cell (when being discharged) for example. Spontaneous reactions (which are in voltaic cells) can be endo or exothermic.

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009 5.0 points

What is the cathode in

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

$Ag^+ + e^- \rightarrow Ag \qquad \mathcal{E}^\circ_{red} = +0.80$
$Fe^{2+} + 2\,e^- \rightarrow Fe \qquad \mathcal{E}^\circ_{red} = -0.44$
and what type cell is it?

1. Ag(s); a battery

2. Fe(s); a battery

3. Not enough information is provided.

4. Ag(s); an electrolytic cell

5. Fe(s); an electrolytic cell **correct**

Explanation:

The diagram A | B || C | D is read as follows:
$A \to B + n\,e^-$ (oxidation)
$C + m\,e^- \to D$ (reduction)
Since reduction occurs at the cathode, the cathode is Fe(s).
To determine the cell type, calculate $\mathcal{E}^\circ$ cell:

$$2\,\text{Ag(s)} \to 2\,\text{Ag}^+\text{(aq)} + 2\,e^-$$
$$\mathcal{E}^\circ_{\text{anode}} = -0.80\text{ V}$$
$$\text{Fe}^{2+} + 2\,e^- \to \text{Fe}$$
$$\mathcal{E}^\circ_{\text{cathode}} = -0.44\text{ V}$$
$$2\,\text{Ag(s)} + \text{Fe}^{2+} \to 2\,\text{Ag}^+\text{(aq)} + \text{Fe}$$
$$\mathcal{E}^\circ_{\text{cell}} = -1.24\text{ V}$$

Since $\mathcal{E}^\circ$ cell is negative, the reaction is spontaneous; *i.e.*, an electrolytic cell.

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010 5.0 points

Standard reduction potentials are established by comparison to the potential of which half reaction?

1. $Li^+ + e^- \longrightarrow Li$

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

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

4. $2\,H_2O + 2\,e^- \longrightarrow H_2 + 2\,OH^-$

5. $2\,H^+ + 2\,e^- \longrightarrow H_2$ **correct**

Explanation:
The hydrogen electrode is the standard reference electrode

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011 5.0 points

If the reduction of fluorine gas to fluoride ion were used as the reference electrode for determining the standard reduction potentials of half reactions

$$F_2(g) + 2\,e^- \to 2\,F^-$$
$$2\,H^+ + 2\,e^- \to H_2(g)$$

what would be the standard reduction potential of the hydrogen electrode?

1. 2.87 V

2. −2.87 V **correct**

3. 3.04 V

4. 0 C

5. −3.04 V

Explanation:
Using the fluorine electrode instead of the hydrogen electrode, *i.e.*, defining the fluorine electrode as zero, would result in the standard reduction potentials of all other half reactions decreasing by 2.87 volts. The hydrogen electrode normally having a standard reduction potential of 0 volts would then have −2.87 volts instead.

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012 5.0 points

The standard potential of the Cu^{2+} | Cu electrode is +0.34 V and the standard potential of the cell

$$\text{Pb(s)}\,|\,\text{Pb}^{2+}\text{(aq)}\,||\,\text{Cu}^{2+}\text{(aq)}\,|\,\text{Cu(s)}$$

is +0.47 V. What is the standard potential of the Pb^{2+} | Pb electrode?

1. +0.81 V

2. − 0.13 V **correct**

3. − 0.26 V

4. +0.13 V

5. − 0.81V

Explanation:
E(cell) = E(cathode) - E(anode)
+0.47V = +0.34V - (?V)
? = -0.13V

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013 5.0 points

The Nernst equation contains the reaction quotient Q. For cells involving solids and aqueous solutions, it is not necessary to include the solids because

1. solids are insoluble in water.

2. solids do not undergo oxidation.

3. solid concentrations (activities) are defined as equal to 1. **correct**

4. solids participate but their concentrations are negligible.

5. solids do not participate in the reactions.

Explanation:

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014 5.0 points

What is the cell potential of the following cell?

$$\text{Pt(s)}\,|\,\text{Br}^-(\text{aq, } 0.2\text{ M})\,|\,\text{Br}_2(\ell)\,||\,\text{Au}^+(\text{aq, } 0.7\text{ M})\,|\,\text{Au(s)}$$

$Br_2 + 2\,e^- \to 2\,Br^-$ $\quad \mathcal{E}^\circ_{\text{red}} = +1.09$ V

$Au^+ + e^- \to Au$ $\quad \mathcal{E}^\circ_{\text{red}} = +1.69$ V

1. 0.55 V **correct**

2. 0.48 V

3. 0.60 V

4. 0.65 V

5. 0.72 V

Explanation:

The reaction involving Br must be reversed (cathode reaction) and the reaction involving Au must be doubled in order to balance the electrons:

$$2\,\text{Br}^-(\text{aq}) \to \text{Br}_2(\ell) + 2\,e^- \qquad \mathcal{E}^\circ_{\text{anode}} = -1.09\text{ V}$$

$$2\,\text{Au}^+(\text{aq}) + 2\,e^- \to 2\,\text{Au(s)} \qquad \mathcal{E}^\circ_{\text{cathode}} = +1.69\text{ V}$$

$$2\,\text{Br}^-(\text{aq}) + 2\,\text{Au}^+(\text{aq}) \to \text{Br}_2(\ell) + 2\,\text{Au(s)} \qquad \mathcal{E}_{\text{cell}} = +0.60\text{ V}$$

To correct for the non-standard concentration we use the Nernst equation:

$$\begin{aligned} E &= E_0 - \frac{0.0591}{n}\log\left(\frac{1}{[\text{Br}^-]^2\,[\text{Au}^+]^2}\right) \\ &= 0.6\text{ V} - \frac{0.0591}{2}\log\frac{1}{(0.2\text{ M})^2\,(0.7\text{ M})^2} \\ &= 0.549536\,V. \end{aligned}$$

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015 5.0 points

A battery formed from the two half reactions

$Fe^{2+} \longrightarrow Fe$ $\quad E^\circ = -0.44$ V

$Cd^{2+} \longrightarrow Cd$ $\quad E^\circ = -0.40$ V

dies (reaches equilibrium). If $[Fe^{2+}]$ was 0.24 M in the dead battery, find $[Cd^{2+}]$ in the dead battery?.

1. 0.0.0005 M

2. 5.4 M

3. 120.3 M

4. 0.01 M **correct**

Explanation:

$E^\circ_{\text{cell}} = +0.04$ V

$$\begin{aligned} E_{\text{cell}} &= E^\circ_{\text{cell}} - \frac{0.05916}{N_e}\log Q \\ 0 &= 0.04 - \frac{0.05916}{2}\log\frac{0.24}{[\text{Cd}^{2+}]} \end{aligned}$$

$$\log\frac{0.24}{[\text{Cd}^{2+}]} = 1.35$$

$$\frac{0.24}{[\text{Cd}^{2+}]} = 10^{1.35}$$

$$[\text{Cd}^{2+}] = \frac{0.24}{10^{1.35}} = 0.0107\text{ M}$$

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016 5.0 points

Sodium is produced by electrolysis of molten sodium chloride. What are the products at the anode and cathode, respectively?

1. $Cl_2(g)$ and $Na_2O(\ell)$

2. $O_2(g)$ and $Na(\ell)$

3. $Na(\ell)$ and $O_2(g)$

4. $Cl_2(g)$ and $Na(\ell)$ **correct**

5. $Cl^-(aq)$ and $Na_2O(\ell)$

Explanation:

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017 5.0 points

The reaction

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

taking place in a battery generates a current of 2 amp. How much Fe(s) is consumed in 1 hour?

1. 1.04 g

2. 4.16 g

3. 8.32 g

4. 3.46 g

5. 2.08 g **correct**

Explanation:

$i = 2$ A $\qquad t = 1$ h

The half equation of interest is

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

and the total charge is

$$q = (2\text{ A})\,(1\text{ h})\,\frac{60\text{ min}}{1\text{ h}}\,\frac{60\text{ sec}}{1\text{ min}} = 7200\text{ C}\,.$$

We can then convert this charge to number of electrons and finally the amount of Fe consumed:

$$(7200\text{ C})\,\frac{1\text{ mol } e^-}{96485\text{ C}} \times \frac{1\text{ mol Fe}}{2\text{ mol } e^-} \times \frac{55.847\text{ g Fe}}{1\text{ mol Fe}} = 2.08374\text{ g Fe}$$

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018 5.0 points

The standard potential of the cell

$Pb(s)\,|\,PbSO_4(s)\,|\,SO_4^{2-}(aq)\,||\,Pb^{2+}(aq)\,|\,Pb(s)$

is +0.23 V at 25°C. Calculate the equilibrium constant for the reaction of 1 M $Pb^{2+}(aq)$ with 1 M $SO_4^{2-}(aq)$.

1. 7.7×10^3

2. 1.7×10^{-8}

3. 6.0×10^7 **correct**

4. 8.0×10^{17}

5. 3.7×10^{16}

Explanation:

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019 5.0 points

You are trying to build a battery, but all you have to work with are some beakers, some wire, a few pieces of Fe(s) and a solution of Fe^{3+}. Is it possible to build a battery? If so, how.

1. Yes, it can be done by preparing two identical half cells.

2. Yes, it can be done by preparing two half cells with very different concentrations of Fe^{3+}. **correct**

3. No, it is not possible.

4. Yes, it can be done by preparing two half cells with very different masses of Fe.

Explanation:

Concentration cells are a practical application of the Nernst equation. By inspection, the Nernst equation reveals that even a reaction for E° is 0, can have a value of E that is greater than zero (i.e. a battery), as long as Q is less than one. This can be accomplished by having very different concentrations of Fe^{3+}, and the battery will run until they are (approximately) equal in both half cells. Note that with such a cell, which half acts as the anode or cathode depends on the concentration gradient; the side with the higher Fe^{3+} concentration will act as the cathode.

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020 5.0 points

Which type of widely used battery is NOT rechargeable?

1. alkaline **correct**

2. lead-acid (storage batteries)

3. nickel-cadmium (NiCad)

4. lithium-ion

Explanation:
Alkaline batteries were not designed to be rechargeable and do not do so efficiently, although there are some websites that disagree.

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021 5.0 points

What condition must be met for a battery to be rechargeable?

1. The electrochemical reaction of the battery must be reversible. **correct**

2. Either its anode or its cathode must generate a gas as a result of the electrochemical reaction.

3. The battery must be open to the outside so that it can vent any internal pressure that builds up from gases within it.

4. It must generate electricity via an acid-base reaction rather than via an oxidation reduction reaction.

Explanation:
Definition

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022 5.0 points

The element Mg is used as a ___?___ for underground pipes by connecting a piece of Mg metal to the pipe with a wire. This method works because Mg is ___?___ .

1. sacrificial electrode; easy to oxidize **correct**

2. protective coating; chemically unreactive

3. sacrificial electrode; easy to reduce

4. protective coating; easy to oxidize

5. protective coating; easy to reduce

Explanation:
Metal underground pipes were made of Pb or Fe. Mg is easier to oxidize (it has a more positive oxidation E) than Pb and Fe, so it makes a good sacrificial anode.

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023 5.0 points

A fuel cell is ___?___ and the corresponding chemical reaction is ___?___ .

1. a voltaic cell; a fusion reaction

2. an electrolytic cell; a combustion reaction

3. an electrolytic cell; a substitution reaction

4. a voltaic cell; a combustion reaction **correct**

5. a voltaic cell; an inihilation reaction

Explanation:

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024 5.0 points

Which of the following is/are desirable when scientists are developing new types of batteries?

I) made from cheap, abundant materials
II) have higher energy density
III) are non-toxic and can be safely recycled

1. I only

2. III only

3. II and III

4. II only

5. I and III

6. I and II

7. I, II and III **correct**

Explanation:
Scientists are continually working to develop new battery technologies. The battery of the future will hopefully be made from cheap, safe materials and will have higher en-

ergy density than the batteries of today.

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025 5.0 points

A photovoltaic cell

1. uses electricity to emit light.

2. converts the energy of a photon into electrical potential energy. **correct**

3. is prohibitively expensive to build and will probably never be commercially available to the general public.

4. is used to plate (deposit) inert metals like gold and silver onto less expensive metals like copper and zinc.

5. is responsible for the functioning of a lead storage battery.

Explanation:

Definition: A photovoltaic cell converts electromagnetic (light) energy into electrical potential energy.