Introduction to Electrolysis

Name____ Section___________

1. Complete as many cells as you can from the data collected in the BCE for today's class. Find other students, who did different experiments, to complete the remaining entries.

Exp	Current (amps)	Time (s)	Mass (g) Zn deposited	Mass (g) Fe deposited	Mass (g) Ag deposited
1	3.00	300		0.26	
2	3.00	600		0.52	
3	2.00	600		0.35	
4	3.00	300	0.30		
5	3.00	600	0.61		
6	2.00	600	0.41		
7	3.00	300			1.01
8	3.00	600			2.01
9	2.00	600			1.34

2. For each metal (Zn, Fe and Ag), how does the mass plated out at the – electrode depend on time? What data from the table in Q1 supports your claim? Write a proportionality equation using the variables mass and time.

There is a direct relationship between the mass of the metal plated out and time as seen in Exp 1 and 2 (Fe), Exp 4 and 5 (Zn) and Exp 7 and 8 (Ag). Note in each pair of experiments for the same metal as the time doubles, the mass plated out also doubles.

Mass α time (seconds)

3. For each metal (Zn, Fe and Ag), how does the mass plated out at the – electrode depend on current? What data from the table in Q1 supports your claim? Write a proportionality equation using the variables mass and current.

There is a direct relationship between the mass of the metal plated out and time as seen in Exp 3 and 2 (Fe), Exp 6 and 5 (Zn) and Exp 9 and 8 (Ag). Note in each pair of experiments for the same metal as the current increases by a factor of 1.5, the mass plated out also increases by a factor of 1.5.

Mass α current (amps)

4. How does the mass of the metal plated out at the – electrode depend on metal (Zn, Fe and Ag)? What data from the table in Q1 supports your claim? Describe any relationship between the metal and mass plated out..

Looking at Exp 1, 4 and 7 (as well as 2, 5 and 8 and 3, 6 and 9) as the time and current are held constant the mass plated out at the cathode always increased from Fe to Zn to Ag.

5. Describe the action that causes the metal to plate on to the – electrode (Note: look at in particulate level animation in the Electrolysis Simulation.). Write a chemical equation (half-reaction) that summarizes your explanation for each of the metals: Zn, Fe and Ag. Summarize similarities and difference among the three half-reactions.

Half-reaction: $Fe^{2+} + 2e^{-} \rightarrow Fe$
Half-reaction: $Zn^{2+} + 2e^{-} \rightarrow Zn$
Half-reaction: $Ag^{+} + 1e^{-} \rightarrow Ag$

If current is defined as the number of electrons passing through a circuit per unit time, and if the current is the same why do we obtain different masses of metal.

Because the molar mass of each metal is different. The half-reactions say that if the same amount of current passes through the electrochemical cell that the moles of Fe and Zn plated would be the same but the moles of Ag would be twice as much.

6a. How many zinc atoms plated out on the – electrode for the three experiments for which the time and current were different for zinc in the table in Q1.

For each experiment:

$$\text{Atoms of Zn} = \text{grams of Zn} \left(\frac{1 \text{ mol}}{65.4 \text{ g}}\right)\left(\frac{6.02 \times 10^{23}}{1 \text{ mol}}\right)$$

$$\text{Exp \#4: Atoms of Zn} = 0.30 \text{ g Zn} \left(\frac{1 \text{ mol}}{65.4 \text{ g}}\right)\left(\frac{6.02 \times 10^{23}}{1 \text{ mol}}\right) = 2.8 \times 10^{21} \text{ atoms}$$

$$\text{Exp \#5: Atoms of Zn} = 0.61 \text{ g Zn} \left(\frac{1 \text{ mol}}{65.4 \text{ g}}\right)\left(\frac{6.02 \times 10^{23}}{1 \text{ mol}}\right) = 5.6 \times 10^{21} \text{ atoms}$$

$$\text{Exp \#6: Atoms of Zn} = 0.41 \text{ g Zn} \left(\frac{1 \text{ mol}}{65.4 \text{ g}}\right)\left(\frac{6.02 \times 10^{23}}{1 \text{ mol}}\right) = 3.8 \times 10^{21} \text{ atoms}$$

b) How many iron atoms plated out on the – electrode for the three experiments for which the time and current were different.

For each experiment:

$$\text{Atoms of Fe} = \text{grams of Fe} \left(\frac{1 \text{ mol}}{55.85 \text{ g}}\right)\left(\frac{6.02 \times 10^{23}}{1 \text{ mol}}\right)$$

$$\text{Exp \#1: Atoms of Fe} = 0.26 \text{ g Fe} \left(\frac{1 \text{ mol}}{55.85 \text{ g}}\right)\left(\frac{6.02 \times 10^{23}}{1 \text{ mol}}\right) = 2.8 \times 10^{21} \text{ atoms}$$

$$\text{Exp \#2: Atoms of Fe} = 0.52 \text{ g Fe} \left(\frac{1 \text{ mol}}{55.85 \text{ g}}\right)\left(\frac{6.02 \times 10^{23}}{1 \text{ mol}}\right) = 5.6 \times 10^{21} \text{ atoms}$$

Exp #3: Atoms of Fe = 0.35 g Fe $\left(\frac{\textbf{1 mol}}{\textbf{55.85 g}}\right)\left(\frac{\textbf{6.02 x 10}^{23}}{\textbf{1 mol}}\right)$ **= 3.8 x 10**^{21} **atoms**

c) How many silver atoms plated out on the – electrode for the three experiments for which the time and current were different.

For each experiment:

Atoms of Ag = grams of Fe $\left(\frac{\textbf{1 mol}}{\textbf{107.9 g}}\right)\left(\frac{\textbf{6.02 x 10}^{23}}{\textbf{1 mol}}\right)$

Exp #7: Atoms of Ag = 1.01 g Ag $\left(\frac{\textbf{1 mol}}{\textbf{107.9 g}}\right)\left(\frac{\textbf{6.02 x 10}^{23}}{\textbf{1 mol}}\right)$ **= 5.6 x 10**^{21} **atoms**

Exp #8: Atoms of Ag = 2.01 g Ag $\left(\frac{\textbf{1 mol}}{\textbf{107.9 g}}\right)\left(\frac{\textbf{6.02 x 10}^{23}}{\textbf{1 mol}}\right)$ **= 11.2 x 10**^{21} **atoms**

Exp #9: Atoms of Ag = 1.34 g Ag $\left(\frac{\textbf{1 mol}}{\textbf{107.9 g}}\right)\left(\frac{\textbf{6.02 x 10}^{23}}{\textbf{1 mol}}\right)$ **= 7.5 x 10**^{21} **atoms**

7. Complete the following table

Exp	Current (amps)	Time (s)	# atoms Zn deposited	# atoms Fe deposited	# atoms Ag deposited
1	3.00	300		0.26	2.8 x 10^{21}
2	3.00	600		0.52	5.6 x 10^{21}
3	2.00	600		0.35	3.8 x 10^{21}
4	3.00	300	0.30		2.8 x 10^{21}
5	3.00	600	0.61		5.6 x 10^{21}
6	2.00	600	0.41		3.8 x 10^{21}
7	3.00	300			5.6 x 10^{21}
8	3.00	600			11.2 x 10^{21}
9	2.00	600			7.5 x 10^{21}

8. What new relationship is made evident by these calculations?

In the three experiments for Zn and Fe the number of atoms of metal plated out is the same. The only thing that is the same for these two metals is the number of electrons transferred. For the experiments involving Ag, twice as many atoms are plated out. So if the number of electrons transferred is half as much for each atom of Ag compared to an atom of Zn or Fe, that could explain the factor of two greater number of atoms plated out.

The new relationship indicates;

$$\textbf{Moles of metal}\ \alpha\ \frac{\textbf{1}}{\textbf{\# electrons transferred}}$$

9. Write the proportionality between the amount of metal (plated out at the – electrode) and each of the three variables (current, time and ion charge). Indicate what variables are held constant for each proportionality.

Moles of metal α time (seconds) when amps and metal are constant

Moles of metal α current (amps) when time and metal are constant

Moles of metal α $\frac{1}{\text{\# electrons transferred}}$ when time and amps are constant

10. Write a single proportionality between the amount of metal (plated out at the – electrode) and the three variables (current, time and ion charge).

Moles of metal α $\frac{\text{time (s)} \cdot \text{amps}}{\text{\# electrons transferred}}$

11. Calculate the magnitude of the constant that will convert the proportionality to an equality.

From Exp #1 (Fe)

Moles of metal α $\frac{\text{time (s)} \cdot \text{amps}}{\text{\# electrons transferred}}$

Constant = $\frac{\text{time (s)} \cdot \text{amps}}{\text{\# electrons transferred moles of metal}}$

Constant = $\frac{300 \text{ s} \cdot 3 \text{ amps}}{2 \cdot 4.7 \times 10^{-3} \text{ mol}} = 9.7 \times 10^{4} \frac{\text{C}}{\text{mol}}$

From Exp #5 (Zn)

Moles of metal α $\frac{\text{time (s)} \cdot \text{amps}}{\text{\# electrons transferred}}$

Constant = $\frac{\text{time (s)} \cdot \text{amps}}{\text{\# electrons transferred moles of metal}}$

Constant = $\frac{600 \text{ s} \cdot 3 \text{ amps}}{2 \cdot 9.3 \times 10^{-3} \text{ mol}} = 9.7 \times 10^{4} \frac{\text{C}}{\text{mol}}$

From Exp #9 (Ag)

$$\textbf{Moles of metal } \alpha \frac{\textbf{time (s)} \cdot \textbf{amps}}{\textbf{\# electrons transferred}}$$

$$\textbf{Constant} = \frac{\textbf{time (s)} \cdot \textbf{amps}}{\textbf{\# electrons transferred moles of metal}}$$

$$\textbf{Constant} = \frac{600 \text{ s} \cdot 2 \text{ amps}}{1 \cdot 1.2 \times 10^{-2} \text{ mol}} = 1.0 \times 10^5 \frac{\text{C}}{\text{mol}}$$

Given the relationship between current (amps), time (seconds) and the charge on the metal ion,

1. Based on the simulation used for today's BCE, describe the change in amount of metal on each electrode. How are these changes related?

The amount of metal that is plated out at the cathode depends on the number of amps passed through the cell, the time the electrons flow, the charge on the metal and the nature of the metal itself (its molar mass).

2. What flows from the + electrode in the external circuit via the wire?

electrons

3. What causes the direction of the flow?

The battery is pumping electrons from the positive electrode to the negative electrode.

4. What flows from the + electrode in the solution?

Metal ions; as electrons are removed from the metal atoms on the electrode metal ions enter the solution.

5. What is the process that transfers the amount of metal from one electrode to the other?

Every time a metal ion enters the solution at the anode, a metal ions in the solution plates out on the cathode. SO the number of metal ions in the solution remain constant.

7. Describe the action that causes the metal atoms to leave the + electrode. Write a chemical equation that summarizes your explanation.

The half-reaction at the anode is:

Half-reaction: $Fe \rightarrow Fe^{2+} + 2e^-$
Half-reaction: $Zn \rightarrow Zn^{2+} + 2e^-$
Half-reaction: $Ag \rightarrow Ag^{+} + e^-$

The half-reaction at the cathode is:
Half-reaction: $Fe^{2+} + 2e^- \rightarrow Fe$
Half-reaction: $Zn^{2+} + 2e^- \rightarrow Zn$
Half-reaction: $Ag^{+} + e^- \rightarrow Ag$