

CHEM*1050 - Experiment 3: Voltaic Cells

Part A. Qualitative Tests of Reducing Strengths of Zn, Pb, Cu and Fe

Using the observations from below, arrange the four **metals** (not ions) in order of **decreasing** reducing strength. Give your reasoning which must follow the experimental results.

Based on my observations, there were 6 reactions giving definite detectable observations:

- 1) Zn/Cu²⁺: Since Zn(s) can reduce Cu²⁺(aq) to Cu(s), Zn(s) has a greater reducing strength than Cu(s).
- 2) Zn/Pb²⁺: Since Zn(s) can reduce Pb²⁺(aq) to Pb(s), Zn(s) has a greater reducing strength than Pb(s).
- 3) Zn/Fe²⁺: Since Zn(s) can reduce Fe²⁺(aq) to Fe(s), Zn(s) has a greater reducing strength than Fe(s).
- 4) Pb/Cu²⁺: Since Pb(s) can reduce Cu²⁺(aq) to Cu(s), Pb(s) has a greater reducing strength than Cu(s).
- 5) Fe/Cu²⁺: Since Fe(s) can reduce Cu²⁺(aq) to Cu(s), Fe(s) has a greater reducing strength than Cu(s).
- 6) Fe/Pb²⁺: Since Fe(s) can reduce Pb²⁺(aq) to Pb(s), Fe(s) has a greater reducing strength than Pb(s).

Zn has the highest reducing strength because it reduced all three of the ions (Cu²⁺, Pb²⁺ and Fe²⁺). Fe reduced the ions Cu²⁺ and Pb²⁺, Pb only reduced Cu²⁺ ions and Cu reduced none of the ions, becoming the metal with the lowest reducing strength.

According to these observations, the order of decreasing reducing strength of these metals is:
Zn(s) > Fe(s) > Pb(s) > Cu(s).

There were no reactions observed between Cu/Zn²⁺, Cu/Pb²⁺, Cu/Fe²⁺, Pb/Zn²⁺, Pb/Fe²⁺ and Fe/Zn²⁺, therefore no conclusions for them.

Observations

Observations		Save & Continue
Did a reaction occur between Cu and Zn²⁺ ?	<input type="radio"/> Yes <input checked="" type="radio"/> No <input type="radio"/> Unsure	
If a reaction occurred write your observations.	No reaction occurred.	
A short but complete answer is ideal.		
If a reaction occurred write a balanced net ionic equation.	No reaction occurred.	
A short but complete answer is ideal.		
If a reaction occurred indicate which metal is the stronger reducing agent. Such as Co > Ag	No reaction occurred.	
A short but complete answer is ideal.		
Did a reaction occur between Cu and Pb²⁺ ?	<input type="radio"/> Yes <input checked="" type="radio"/> No <input type="radio"/> Unsure	
If a reaction occurred write your observations.	No reaction occurred.	
A short but complete answer is ideal.		
If a reaction occurred write a balanced net ionic equation	No reaction occurred.	
A short but complete answer is ideal.		
If a reaction occurred indicate which metal is the stronger reducing agent.	No reaction occurred.	
A short but complete answer is ideal.		

Did a reaction occur between
Zn and Pb²⁺ ?

Yes No Unsure

If a reaction occurred write your
observations.

Initial: Shiny and metallic metal, grey in colour.
During: Zn metal flakes observed in solution.

A short but complete answer is ideal.

If a reaction occurred write a
balanced net ionic equation

$$\text{Zn(s)} + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Pb(s)}$$

A short but complete answer is ideal.

If a reaction occurred indicate which
metal is the stronger reducing
agent.

$$\text{Zn(s)} > \text{Pb(s)}$$

A short but complete answer is ideal.

Did a reaction occur between
Zn and Fe²⁺ ?

Yes No Unsure

If a reaction occurred write your
observations.

Initial: Shiny and metallic metal, grey in colour.
During: Zn metal flakes observed in solution.

A short but complete answer is ideal.

If a reaction occurred write a
balanced net ionic equation

$$\text{Zn(s)} + \text{Fe}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Fe(s)}$$

A short but complete answer is ideal.

If a reaction occurred indicate which
metal is the stronger reducing
agent.

$$\text{Zn(s)} > \text{Fe(s)}$$

A short but complete answer is ideal.

Did a reaction occur between
Pb and Cu²⁺ ?

Yes No Unsure

If a reaction occurred write your
observations.

Initial: Shiny and metallic metal, grey in colour.
During: Pb metal flakes observed in solution. Metal

A short but complete answer is ideal.

If a reaction occurred write a
balanced net ionic equation

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

A short but complete answer is ideal.

If a reaction occurred indicate which
metal is the stronger reducing
agent.

$$\text{Pb(s)} > \text{Cu(s)}$$

A short but complete answer is ideal.

Did a reaction occur between
Pb and Zn²⁺ ?

Yes No Unsure

If a reaction occurred write your
observations.

No reaction occurred.

A short but complete answer is ideal.

If a reaction occurred write a
balanced net ionic equation

No reaction occurred.

A short but complete answer is ideal.

If a reaction occurred indicate which
metal is the stronger reducing
agent.

No reaction occurred.

A short but complete answer is ideal.

Did a reaction occur between
Pb and Fe²⁺ ?

Yes No Unsure

If a reaction occurred write your
observations.

No reaction occurred.

A short but complete answer is ideal.



If a reaction occurred write a
balanced net ionic equation

No reaction occurred.

A short but complete answer is ideal.



If a reaction occurred indicate which
metal is the stronger reducing
agent.

No reaction occurred.

A short but complete answer is ideal.



Did a reaction occur between
Fe and Cu²⁺ ?

Yes No Unsure

If a reaction occurred write your
observations.

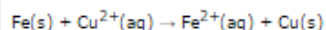
Initial: Dull metal is grey in colour.

During: Fe metal starts to change colour.

A short but complete answer is ideal.



If a reaction occurred write a
balanced net ionic equation



A short but complete answer is ideal.



If a reaction occurred indicate which
metal is the stronger reducing
agent.

Fe(s) > Cu(s)

A short but complete answer is ideal.



Did a reaction occur between
Fe and Zn²⁺ ?

Yes No Unsure

If a reaction occurred write your
observations.

No reaction occurred.

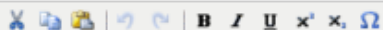
A short but complete answer is ideal.



If a reaction occurred write a
balanced net ionic equation

No reaction occurred.

A short but complete answer is ideal.



If a reaction occurred indicate which
metal is the stronger reducing
agent.

No reaction occurred.

A short but complete answer is ideal.



Did a reaction occur between
Fe and Pb²⁺ ?

Yes No Unsure

If a reaction occurred write your
observations.

Initial: Dull metal is grey in colour

During: Metal starts to change colour.

A short but complete answer is ideal.



If a reaction occurred write a
balanced net ionic equation



A short but complete answer is ideal.



If a reaction occurred indicate which
metal is the stronger reducing
agent.

Fe(s) > Pb(s)

A short but complete answer is ideal.



Did a reaction occur between **Cu and Fe²⁺** ? Yes No Unsure

If a reaction occurred write your observations.

A short but complete answer is ideal.

No reaction occurred.

✂ 📄 🗑 ↶ ↷ | **B I U x² x₂ Ω**

If a reaction occurred write a balanced net ionic equation

A short but complete answer is ideal.

No reaction occurred.

✂ 📄 🗑 ↶ ↷ | **B I U x² x₂ Ω**

If a reaction occurred indicate which metal is the stronger reducing agent.

A short but complete answer is ideal.

No reaction occurred.

✂ 📄 🗑 ↶ ↷ | **B I U x² x₂ Ω**

Did a reaction occur between **Zn and Cu²⁺** ? Yes No Unsure

If a reaction occurred write your observations.

A short but complete answer is ideal.

Initial: Shiny and metallic metal, grey in colour.
During: Zn metal flakes observed in solution.

✂ 📄 🗑 ↶ ↷ | **B I U x² x₂ Ω**

If a reaction occurred write a balanced net ionic equation

A short but complete answer is ideal.

$$\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$$

✂ 📄 🗑 ↶ ↷ | **B I U x² x₂ Ω**

If a reaction occurred indicate which metal is the stronger reducing agent.

A short but complete answer is ideal.

$$\text{Zn(s)} > \text{Cu(s)}$$

✂ 📄 🗑 ↶ ↷ | **B I U x² x₂ Ω**

Zn and Cu²⁺

Initial: Shiny and metallic metal, grey in colour.

During: Zn metal flakes observed in solution.

Final: Black precipitate formed on the metal.

Zn and Pb²⁺

Initial: Shiny and metallic metal, grey in colour.

During: Zn metal flakes observed in solution.

Final: Black precipitate formed on the metal.

Zn and Fe²⁺

Initial: Shiny and metallic metal, grey in colour.

During: Zn metal flakes observed in solution.

Final: Black precipitate formed on the metal.

Pb and Cu²⁺

Initial: Shiny and metallic metal, grey in colour.

During: Pb metal flakes observed in solution.

Final: Black precipitate formed on the metal.

Fe and Cu²⁺

Initial: Dull metal is grey in colour.

During: Fe metal starts to change colour.

Final: Copper coloured precipitate formed on the metal.

Fe and Pb²⁺

Initial: Dull metal is grey in colour

During: Metal starts to change colour.

Final: Dark grey precipitate formed on the metal.

Part B. Measurement of Cell Potentials

Save&Continue

Measurements and calculations should be entered as a number followed by units. Nothing else. Remember to report to the correct precision or significant figures as required.

Voltage of Zn-Cu cell (1 st determination)	<input type="text" value="1.076 V"/>	
Voltage of Zn-Cu cell (2 nd determination)	<input type="text" value="1.076 V"/>	
Average cell voltage	<input type="text" value="1.076 V"/>	
At which metal did reduction occur?	<input type="text" value="Cu"/>	Zn or Cu
At which metal did oxidation occur?	<input type="text" value="Zn"/>	
Which metal is the stronger reducing agent?	<input type="text" value="Zn"/>	Zn or Cu
Theoretical Voltage (assume 0.5M)	<input type="text" value="1.10 V"/>	
Voltage of Cu-Pb cell (1 st determination)	<input type="text" value="0.458 V"/>	
Voltage of Cu-Pb cell (2 nd determination)	<input type="text" value="0.462 V"/>	
Average cell voltage	<input type="text" value="0.460 V"/>	
At which metal did reduction occur?	<input type="text" value="Cu"/>	Cu or Pb
At which metal did oxidation occur?	<input type="text" value="Pb"/>	
Which metal is the stronger reducing agent?	<input type="text" value="Pb"/>	Cu or Pb
Theoretical Voltage (assume 0.5M)	<input type="text" value="0.47 V"/>	
Voltage of Cu-Fe cell (1 st determination)	<input type="text" value="0.555 V"/>	
Voltage of Cu-Fe cell (2 nd determination)	<input type="text" value="0.578 V"/>	
Average cell voltage	<input type="text" value="0.567 V"/>	
At which metal did reduction occur?	<input type="text" value="Cu"/>	Cu or Fe
At which metal did oxidation occur?	<input type="text" value="Fe"/>	
Which metal is the stronger reducing agent?	<input type="text" value="Fe"/>	Cu or Fe
Theoretical Voltage (assume 0.5M)	<input type="text" value="0.78 V"/>	
Voltage of Zn-Pb cell (1 st determination)	<input type="text" value="0.593 V"/>	
Voltage of Zn-Pb cell (2 nd determination)	<input type="text" value="0.593 V"/>	
Average cell voltage	<input type="text" value="0.593 V"/>	
At which metal did reduction occur?	<input type="text" value="Pb"/>	Zn or Pb
At which metal did oxidation occur?	<input type="text" value="Zn"/>	
Which metal is the stronger reducing agent?	<input type="text" value="Zn"/>	Zn or Pb
Theoretical Voltage (assume 0.5M)	<input type="text" value="0.63 V"/>	

Voltage of **Fe-Zn** cell (1st determination)

Voltage of **Fe-Zn** cell (2nd determination)

Average cell voltage

At which metal did reduction occur? Fe or Zn

At which metal did oxidation occur?

Which metal is the stronger reducing agent? Fe or Zn

Theoretical Voltage (assume 0.5M)

Voltage of **Fe-Pb** cell (1st determination)

Voltage of **Fe-Pb** cell (2nd determination)

Average cell voltage

At which metal did reduction occur? Fe or Pb

At which metal did oxidation occur?

Which metal is the stronger reducing agent? Fe or Pb

Theoretical Voltage (assume 0.5M)

On the basis of theoretical voltages, write the order for reducing strength of the four metals, in order of decreasing reducing strength

A short but complete answer is ideal.

TA assigned cell #1

Save&Continue

TA assigned cell #1

A short but complete answer is ideal.

Write the anode reaction for the TA assigned cell #1.

A short but complete answer is ideal.

Write the cathode reaction for the TA assigned cell #1.

A short but complete answer is ideal.

TA assigned cell #2

Fe - Zn

For TA assigned cell #2 show how you would calculate the Fe²⁺ solution concentration using the Nernst equation and your measurements from Part B

$$E = 0.481 \text{ V}, \quad E^\circ = 0.32 \text{ V}$$

$$n = 2 \text{ mol}, \quad R = 8.3145 \text{ J/K mol}$$

$$F = 96485 \text{ C/mol of } e^-, \quad T = 298 \text{ K}$$

$$[\text{Zn}^{2+}] = 0.5 \text{ M} \quad [\text{Fe}^{2+}] = ?$$

$$\text{Using Nernst equation } E = E^\circ - (RT/nF)\ln Q$$

rearranged to:

$$Q = e^{((E^\circ - E)nF)/RT}$$

$$[\text{anode}]/[\text{cathode}] = e^{((E^\circ - E)nF)/RT}$$

$$[\text{Zn}^{2+}]/[\text{Fe}^{2+}] = e^{((E^\circ - E)nF)/RT}$$

$$[0.5\text{M}]/[\text{Fe}^{2+}] = e^{(((0.32\text{V} - 0.481\text{V})(2 \text{ mol})(96485 \text{ C/mol}))/((8.314\text{J/K mol})(298\text{K}))}$$

$$[0.5\text{M}]/[\text{Fe}^{2+}] = e^{-12.53976474}$$

$$[\text{Fe}^{2+}] = [0.5 \text{ M}]/e^{-12.53901065}$$

$$= [0.5 \text{ M}]/(3.58407317 \times 10^{-6})$$

$$= 1.395060804 \times 10^5 \text{ M}$$

Reported with 1 sig figs:

$$[\text{Fe}^{2+}] = 1 \times 10^5 \text{ M}$$

Explain whether your answer is reasonable and whether it makes mathematical sense? Does it make chemical sense?

The answer makes sense mathematically because according to the Nernst equation, an increase in cell voltage is caused by an increase in the cathode concentration, in this case the Fe²⁺ concentration. Since the Fe²⁺ concentration is calculated to be more than the Zn²⁺ concentration, it is plausible mathematically. However, it does not make sense chemically, as the calculated concentration of Fe²⁺ is too large to be plausible.

Part C. Influence of Concentration on Cell Voltage

Save&Continue

Measured Temperature of **Zn-Cu** cell

21.8 C

Use C for degrees Celcius or K for Kelvin.

Voltage of **Zn-Cu** cell
(0.005 M Cu²⁺)

1.038 V

Measured Temperature of **Cu-Cu** cell

21.8 C

Voltage of **Cu-Cu** cell
(0.500 M/0.005 M Cu²⁺)

- 0.030 V

Voltage of **Zn-Cu** cell
(0.500 M Cu²⁺ from Part B)

1.076 V

Describe what happened to the measured voltage (Zn-Cu) when the cathode solution was diluted.

Dilution of $\text{Cu}^{2+}(\text{aq})$ from 0.5 M to 0.005 M, makes the voltage of the diluted Zn-Cu cell (1.038 V) lower than the undiluted Zn-Cu cell (1.067 V).

Also, upon dilution, $\ln Q$ becomes larger, causing the cell voltage (E) to decrease. This can be shown as following:

Undiluted

$$\begin{aligned}\ln Q &= \ln\left(\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}\right) \\ &= \ln(0.5\text{M}/0.5\text{M}) = \ln(1) = 0\end{aligned}$$

$$\text{So, } E = E^\circ$$

Diluted

$$\begin{aligned}\ln Q &= \ln\left(\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}\right) \\ &= \ln(0.5\text{M}/0.005\text{M}) = \ln(100)\end{aligned}$$

$$\text{So, } E = E^\circ - \ln(100) = E^\circ - 4.605$$

$$E_{\text{undiluted}} > E_{\text{diluted}} \rightarrow E^\circ > E^\circ - \ln(100)$$

Therefore, the measured voltage (E) decreases when diluted.

What do you expect would have happened if the anode solution had been diluted instead of the cathode solution (Zn-Cu)?

If the anode solution was diluted instead of the cathode solution, the concentration of Zn^{2+} would have decreased. This will also cause the value of Q to become smaller ($1 \rightarrow 0.01$), and since $\ln Q = \ln\left(\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}\right)$, the value of $\ln Q$ decreases, causing the cell voltage (E) to increase.

Undiluted

$$\begin{aligned}\ln Q &= \ln\left(\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}\right) \\ &= \ln(0.5\text{M}/0.5\text{M}) = \ln(1) = 0\end{aligned}$$

$$\text{So, } E = E^\circ$$

Diluted

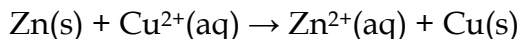
$$\begin{aligned}\ln Q &= \ln\left(\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}\right) \\ &= \ln(0.005\text{M}/0.5\text{M}) = \ln(0.01)\end{aligned}$$

$$\text{So, } E = E^\circ - \ln(0.01) = E^\circ + 4.605$$

$$E_{\text{undiluted}} < E_{\text{diluted}} \rightarrow E^\circ < E^\circ - \ln(0.01)$$

Therefore, diluting the anode solution will cause the measured voltage (E) to increase.

Show how you would calculate the theoretical voltage of the Zn-Cu cell in Part C, using standard reduction potentials and the measured temperature.



$$E^\circ = 0.34 \text{ V} + 0.76 \text{ V} = 1.10 \text{ V}$$

$$n = 2 \text{ mol}$$

$$T = 21.8 \text{ C} + 273 = 294.8 \text{ K}$$

$$[\text{Zn}^{2+}] = 0.5\text{M}$$

$$[\text{Cu}^{2+}] = 0.005\text{M}$$

$$\ln Q = [\text{Zn}^{2+}]/[\text{Cu}^{2+}] = 0.5/0.005 = 100$$

Using Nernst equation, $E = E^\circ - (RT/nF)\ln Q$, the cell voltage (E) for the Zn-Cu cell can be calculated.

$$E = 1.10 \text{ V} - ((8.3145 \text{ J/K mol})(294.8 \text{ K})/((2\text{mol})(96485 \text{ C/mol})) \ln(100))$$

$$E = 1.041504898 \text{ V}$$

Theoretical Voltage of **Zn-Cu** cell
(0.005 M Cu^{2+} from Part C)

$$1.04 \text{ V}$$

To form a voltaic cell, should the voltage be positive or negative?

To form a voltaic cell, the voltage should be positive.

Which compartment in the Cu-Cu cell forms the anode to make a voltaic cell?

Copper (0.005 M) forms the anode to make a voltaic cell.

Show how you would calculate the theoretical voltage of the voltaic Cu-Cu cell in Part C, using standard reduction potentials and the measured temperature.

$$E^\circ = 0$$

$$T = 21.8 \text{ C} + 273 = 294.8 \text{ K}$$

$$Q = [\text{Cu}_{\text{diluted}}]/[\text{Cu}_{\text{undiluted}}]$$
$$= 0.005\text{M}/0.5\text{M} = 0.01$$

$$E = E^\circ - (RT/nF)\ln Q$$

$$E = 0 - ((8.3145 \text{ J/Kmol})(294.8 \text{ K})/((2 \text{ mol})(96485 \text{ C/mol}))\ln(0.01))$$

$$E = 0.058495102 \text{ V}$$

Theoretical Voltage of the **voltaic Cu-Cu** cell
(0.005M/0.5M Cu^{2+} from Part C)

$$0.058 \text{ V}$$