

## Appendix 2 – Detailed practical lessons

### 1. Electrochemical cells and chemical change

#### Student Sheet

In this exercise, you will construct electrochemical cells and measure their e.m.f. values. You will compare your results with the results of test tube reactions, and investigate the effect of changing cell conditions.

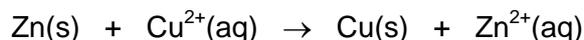
#### Intended lesson outcomes

By the end of this exercise you will be able to

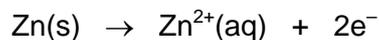
- set up an electrochemical cell
- read a voltmeter
- record observations from test tube experiments
- compare results from different types of experiment
- write cell descriptions and calculate  $E_{\text{cell}}$

#### Background information

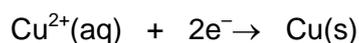
Normally, for two species to react together it is necessary for them to collide with each other. So, when zinc metal is placed in aqueous copper sulphate a redox reaction occurs when the  $\text{Cu}^{2+}$  ions collide with the zinc metal. The redox reaction is exothermic and so heat energy is also evolved. The ionic equation for this reaction is given below.



In this reaction, Zn atoms lose electrons (they are **oxidised**, and so zinc acts as a **reducing agent**).

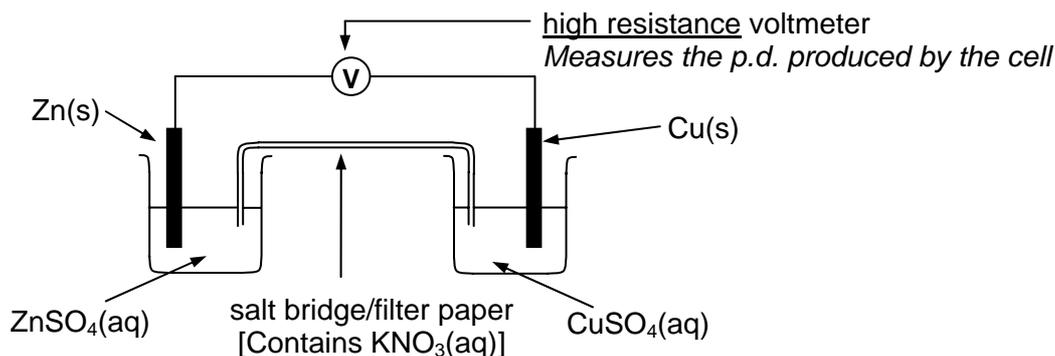


At the same time,  $\text{Cu}^{2+}$  ions gain electrons (they are **reduced**, and so are acting as an **oxidising agent**).



During the collision, electrons are transferred **directly** from zinc atoms to  $\text{Cu}^{2+}$  ions. If zinc metal is used as the electrode in one half of an electrochemical cell, and  $\text{Cu}^{2+}$  ions are present in the other half-cell, exactly the same reaction can occur. However, in order for the electron transfer to take place, electrons travel from the zinc atoms to the  $\text{Cu}^{2+}$  ions **via a metal wire**. In this case, most of the energy released by the reaction is obtained in the form of an electric current flowing in the wire.

The diagram below shows an electrochemical cell between zinc and copper. A strip of zinc foil is placed in a solution containing zinc ions, and a strip of copper is placed in a solution of copper ions. The two metals are joined by wire, and a salt bridge (e.g. a strip of filter paper soaked in saturated potassium nitrate solution) completes the circuit.



The zinc, being the more reactive metal, releases electrons more easily than copper. The electron density on the zinc foil is, therefore, higher than that on the copper foil. Electrons then flow from the area with higher electron density to the area with lower electron density. Thus, the zinc foil is the **negative pole** (relatively higher electron density), and the copper is the **positive pole** (relatively lower electron density) of this cell.

The cell works because of the **difference** in electron densities on each pole. The greater this difference, the larger the potential difference (p.d.) produced. It is, therefore, the **chemistry** of the process that determines the value of the p.d. of a cell. Increasing the **size** of the pieces of foil would increase the ability of the cell to pass electricity, so **allowing a larger current** to flow, but would have **no effect** on the p.d. produced. A cell would produce its maximum p.d. when no current is flowing. This is known as the *electromotive force*, i.e. the e.m.f. of the cell, or the Standard Cell Potential,  $E^\ominus_{\text{cell}}$ .

When working, a cell does not produce this maximum value, so we refer to measure the p.d. of a cell, rather than its e.m.f. value. When **measuring** a p.d. we use a **high resistance voltmeter**, the purpose of which is to reduce the current flow as near to zero as we can. The p.d. measured is then as close to the e.m.f. as we can get.

Instead of drawing each cell out in full, as above, by convention we can write a **Cell Description**. A *cell description* shows on its **left hand side** the chemical change occurring in the half-cell at the **negative pole**. On its **right hand side**, the chemical change occurring in the half-cell at the **positive pole** is shown.

The cell description of the above cell is shown below.



- The **single line** ' | ' placed between the two species indicates that they are in different phases (solid metal and aqueous ions in this case).
- The order of writing the species is: reactant | product.
- The **double line** between the two half-cells represents the **salt bridge**.

### Safety

Materials safety data sheets should be consulted so that the correct action can be taken in event of a spillage and/or accident.

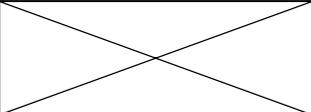
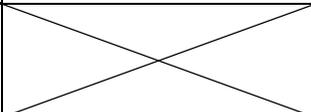
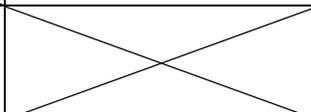
	<b>You must wear eye protection throughout this experiment</b>
	Potassium nitrate is <b>oxidising</b>
	Copper sulphate and zinc sulphate are <b>irritants</b> Iron(II) sulphate and iron(III) chloride are <b>harmful</b>
	Silver nitrate is <b>corrosive</b>
	Copper sulphate, zinc sulphate and silver nitrate are <b>dangerous for the environment</b> . Your teacher will tell you how to dispose of these

### Procedure

**This exercise is designed to show the link between** what could be called **direct redox reactions**, when reactants are mixed together, and the **indirect** redox reactions that occur in an electrochemical cell.

#### Experiment 1

- Clean the pieces of metal with wire wool or sand paper to remove any surface coating. Treat the silver gently and only scour if it is not shiny.
- Perform test tube reactions for each of the metal and salt combinations shown in the table below. After each test examine the surface of the metal to see if there is any change in its appearance, which would suggest that a reaction has taken place.
- Record your observations in the table.

metal/salt	Cu	Zn	Ag
<b>CuSO<sub>4</sub></b>			
<b>ZnSO<sub>4</sub></b>			
<b>AgNO<sub>3</sub></b>			

**Questions on Experiment 1**

For those experiments in which a **change is observed**:

- 1 Write an ionic equation for the overall reaction.
- 2 Write half-equations (ion–electron equations) for the oxidation and the reduction process.
- 3 From these half-equations, deduce what is being oxidised and what is being reduced. Explain your answers.
- 4 Name the oxidising agent and the reducing agent in the reaction.

**Experiment 2**

All solutions used in **this** experiment must have metal ion concentrations of **0.10 mol dm<sup>-3</sup>**.

1. Set up the circuit as shown in the diagram on the previous page.
2. To complete the circuit for each cell, a salt bridge must be used. This can be a strip of filter paper soaked in saturated potassium nitrate solution. Use a **new salt bridge** each time.
3. Construct each of the cells given below and measure the p.d. produced with a voltmeter. You should test the copper/zinc cell **last**, as you will use this cell again in **Experiment 3**.
4. Connect the voltmeter so that it shows a **positive** value. We will call this the cell e.m.f.,  $E_{\text{cell}}$ .
5. **Note:** we do not use the standard symbol ( $E^{\ominus}_{\text{cell}}$ ), as the concentrations used in the half-cells are not 1.00 mol dm<sup>-3</sup>. However, as the concentrations in the half-cells are equal, the value obtained will not be that far away from the standard value.
6. In each case deduce the **polarity** of the cell (by noting which of the electrodes is connected to the **black** terminal of the voltmeter). This is the negative pole of the cell.
7. Record your results in the table below.
8. The cell description for the copper/zinc cell is written in for you. Write a cell description for the other two cells.

	cell description	$E_{\text{cell}} / \text{V}$	negative pole	direction of electron flow
copper/silver cell				from:- to:-
silver/zinc cell				from:- to:-
copper/zinc cell	$\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}) \parallel \text{Cu}^{2+}(\text{aq}) \mid \text{Cu(s)}$			from:- to:-

- Compare the results from Experiment 2 with the results from the equivalent metal/salt combinations in Equation 1.

**Note**

The chemical changes taking place when a cell is producing a current are the same ones that took place in Experiment 1 so the equations will be the same.

**Questions on Experiment 2**

- 1 What can you deduce about the role of the half-cell which forms the negative pole of each cell?
- 2 State what is happening to each metal strip.
- 3 Write half-equations for the reaction in each half-cell.
- 4 Write an overall equation for the reaction occurring in each cell.
- 5 Use the standard electrode potential,  $E^\ominus$  data given below to calculate  $E^\ominus_{\text{cell}}$  values for each cell using the formula:

$$E^\ominus_{\text{cell}} = E^\ominus (\text{RHS}) - E^\ominus (\text{LHS})$$



- 6 Consider the cell:  $\text{Ag}(\text{s}) | \text{Ag}^+(\text{aq}) || \text{Zn}^{2+}(\text{aq}) | \text{Zn}(\text{s})$

Identify the test in Experiment 1 which has the same metal/salt combination as in this cell. Does this test work?

Identify the test in Experiment 1 in which metal/salt combination as in this cell is reversed. Does this test work?

Calculate the  $E_{\text{cell}}$  value for the cell using:  $E^\ominus_{\text{cell}} = E^\ominus (\text{RHS}) - E^\ominus (\text{LHS})$

By reference to this  $E_{\text{cell}}$  value, state whether the reaction in this cell is feasible or non-feasible. Explain your answer.

**Experiment 3**

Using the copper/zinc cell, replace the original copper sulphate solution, of concentration  $0.10 \text{ mol dm}^{-3}$ , with:

- (i) A solution of copper sulphate of concentration  $1.0 \text{ mol dm}^{-3}$
- (ii) A solution of copper sulphate of concentration  $0.010 \text{ mol dm}^{-3}$
- (iii) A solution of copper sulphate of concentration  $0.0010 \text{ mol dm}^{-3}$

In each case, measure and record the  $E_{\text{cell}}$  value obtained. Compare your results with your result from Experiment 2.

**Questions on Experiment 3**

- 1 What effect did diluting the copper sulphate solution have on the  $E^{\ominus}_{\text{cell}}$  value of the copper/zinc cell?
- 2 Using Le Chatelier's principle, state and explain what effect there will be on the electron density on the copper foil when the copper sulphate solution is diluted?
- 3 Using your answer to part 2, account for the variation in  $E^{\ominus}_{\text{cell}}$  value when the copper sulphate solution was diluted.
- 4 Predict, with explanation, the effect of diluting the zinc sulphate solution, while using copper sulphate solution of concentration  $1.0 \text{ mol dm}^{-3}$ .

**Experiment 4**

Prepare a sample of solution **A**, by thoroughly mixing together  $5.0 \text{ cm}^3$  of aqueous  $\text{FeSO}_4$ , of concentration  $0.10 \text{ mol dm}^{-3}$  and  $5.0 \text{ cm}^3$  of aqueous  $\text{FeCl}_3$ , of concentration  $0.10 \text{ mol dm}^{-3}$ .

Perform the following tests.

- To a small amount of a solution of iodine in potassium iodide solution add a few drops of starch solution.
- Divide solution **A** into two portions.
- To the first portion of solution **A** add a few drops of starch solution.
- To the second portion of solution **A** add an equal volume of potassium iodide solution; then add a few drops of starch solution.

**Questions on Experiment 4**

It should be clear from your observations that a redox reaction took place when solution **A** was mixed with potassium iodide solution.

- 1 Identify the species that has been oxidised, and the species that has been reduced.
- 2 Write half-equations for the oxidation process and the reduction process.
- 3 Combine the half-equations to give an overall equation for the reaction.
- 4 Name the oxidising agent and the reducing agent.
- 5 Obtain appropriate standard electrode potential,  $E^{\ominus}$  data values from a data book and use them to calculate the  $E^{\ominus}_{\text{cell}}$  value for this feasible reaction.

## 1. Electrochemical cells and chemical change

### Teachers' Notes

Overall, the exercise provides a range of learning experiences for the student, which include: performing test tube redox reaction, constructing electrochemical cells and comparing the two sets of results obtained. It is possible that these experiments could be performed as a circus experiment. This would reduce the amount of equipment needed for the group.

#### Intended learning outcomes

Please see the Student Sheet.

#### A suggested approach

The aim of this exercise is to test or to develop the understanding students have of the operation of electrochemical cells, the construction of using different electrochemical cells, the effect of changing concentration on the Standard Cell Potential,  $E^{\circ}_{\text{cell}}$ , value and the use of Standard Electrode Potential,  $E^{\circ}$ , values in predicting the feasibility of reactions. You may need to remind your students that although conventionally we refer to current flowing from the positive pole to the negative pole, in reality, the electron flow is from negative to positive.

The experimentation should not take too long, but the students are likely to need a considerable amount of 'thinking time' to answer the questions posed in the exercise.

**Experiments 1 and 2** are linked, as the same redox reactions are involved in both. It would be worth spending a short time reviewing the basic theory involved, perhaps by reference to the Student Sheet.

It would be useful before **Experiment 3** is attempted to review the effect of concentration changes on equilibrium positions. Using Le Chatelier's principle students could be asked to predict the effect dilution would have on the electron densities on the metal foils and hence deduce the likely direction of change in  $E^{\circ}_{\text{cell}}$  values.

**Experiment 4** provides an opportunity for results to be analysed and an explanation for the redox reaction occurring to be deduced.

While the class is working on this exercise, you might like to 'entertain' them by growing a **silver tree!**

Using scissors, cut copper foil roughly into the shape of a tree with many small 'branches' (perhaps a Christmas tree if this is familiar to your students). Place it in a tall beaker and cover it with a well-diluted solution of silver nitrate. Silver crystals grow on the 'branches' and after some time hide the copper completely. The result is a rather pretty silver tree.

#### Answers to questions

##### Experiment 1

- 1 E.g.  $\text{Zn} + 2\text{Ag}^+ \rightarrow \text{Zn}^{2+} + 2\text{Ag}$
- 2 Oxidation:  $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ ; Reduction:  $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$
- 3 Zn oxidised;  $\text{Ag}^+$  reduced
- 4 Oxidising agent = silver nitrate; Reducing agent = zinc sulphate

**Experiment 2**

- 1 It provides electrons so is the reducing agent
- 2 E.g. Zinc strip dissolves; silver is deposited on silver strip
- 3  $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ ;  $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$
- 4  $\text{Zn} + 2\text{Ag}^+ \rightarrow \text{Zn}^{2+} + 2\text{Ag}$
- 5  $E^\ominus_{\text{cell}}(\text{Zn}/\text{Cu}^{2+}) = 1.10 \text{ V}$ ;  $E^\ominus_{\text{cell}}(\text{Zn}/\text{Ag}^+) = 1.56 \text{ V}$ ;  $E^\ominus_{\text{cell}}(\text{Cu}/\text{Ag}^+) = 0.46 \text{ V}$
- 6 Same combination =  $\text{Ag}(\text{s})/\text{Zn}^{2+}$  (doesn't work)  
Reversed combination =  $\text{Zn}(\text{s})/\text{Ag}^+$  (does work)  
 $E^\ominus_{\text{cell}}(\text{Ag}/\text{Zn}^{2+}) = -1.56 \text{ V}$ ; reaction non-feasible as  $E^\ominus_{\text{cell}}$  is negative – reverse reaction feasible

**Experiment 3**

- 1 Diluting  $\text{Cu}^{2+}$  **lowered**  $E_{\text{cell}}(\text{Zn}/\text{Cu}^{2+})$
- 2  $\text{Cu}(\text{s}) \rightleftharpoons \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$ ; Lower  $[\text{Cu}^{2+}]$ ; equilibrium to right; increases  $\text{e}^-$  density on Cu foil
- 3 Difference in  $\text{e}^-$  density between two poles reduced;  $E^\ominus_{\text{cell}}(\text{Zn}/\text{Cu}^{2+})$  falls  
(opposite argument for increased  $[\text{Cu}^{2+}]$ )
- 4  $\text{Zn}(\text{s}) \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ ; Lower  $[\text{Zn}^{2+}]$ ; equilibrium to right; increases  $\text{e}^-$  density on Zn foil  
Difference in  $\text{e}^-$  density between two poles increased;  $E^\ominus_{\text{cell}}(\text{Zn}/\text{Cu}^{2+})$  rises

**Experiment 4**

- 1 Iodide ions oxidised,  $\text{Fe}^{3+}$  ions reduced
- 2 Oxidation:  $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$ ; Reduction:  $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$
- 3  $2\text{Fe}^{3+} + 2\text{I}^- \rightarrow 2\text{Fe}^{2+} + \text{I}_2$
- 4 Oxidising agent = Iron(III) ion/chloride; Reducing agent = potassium iodide / iodide ions
- 5  $E^\ominus_{\text{cell}} = (+0.77) - (+0.54) = +0.23 \text{ V}$

## Technical information

### Requirements per student/group

- **Note:** it is only necessary to make up a sufficient volume of each solution to meet the needs of the class. For example, if the exercise were to be performed as a circus, then 50 cm<sup>3</sup> of silver nitrate solution would be ample. In this case, students should be instructed not to throw the solutions away after Experiment 1 but to pass them on to another group. The metals should be rinsed with water before re-use and any deposit can be removed with wire wool.
- Eye protection
- Six test tubes
- Two 10 cm<sup>3</sup> measuring cylinders
- Distilled water bottle
- Three 100 cm<sup>3</sup> beakers
- Wires and crocodile clips
- Access to a high resistance voltmeter
- Filter paper cut into strips of sufficient length to join the two 100 cm<sup>3</sup> beakers
- Saturated potassium nitrate solution
- Strips of zinc, copper and silver of suitable size
- Wire wool / sand paper
- Two 100 cm<sup>3</sup> conical flasks
- Access to 1.0 mol dm<sup>-3</sup> copper sulphate solution (make by dissolving 250 g of hydrated copper sulphate in distilled water and make up to 1.0 dm<sup>3</sup>)
- 50 cm<sup>3</sup> of 0.10 mol dm<sup>-3</sup> copper sulphate solution (make by dilution)
- 50 cm<sup>3</sup> of 0.010 mol dm<sup>-3</sup> copper sulphate solution (make by dilution)
- 50 cm<sup>3</sup> of 0.0010 mol dm<sup>-3</sup> copper sulphate solution (make by dilution)
- 50 cm<sup>3</sup> of 0.10 mol dm<sup>-3</sup> zinc sulphate solution (make by dissolving 28.8 g of hydrated zinc sulphate in distilled water and make up to 1.0 dm<sup>3</sup>)
- 50 cm<sup>3</sup> of 0.10 mol dm<sup>-3</sup> silver nitrate solution (make by dissolving 16.98 g of silver nitrate (**Corrosive**) in distilled water and make up to 1.0 dm<sup>3</sup>)
- Access to 0.10 mol dm<sup>-3</sup> iron(II) sulphate solution (make by dissolving 27.8 g of hydrated iron(II) sulphate (**Harmful**) in 200 cm<sup>3</sup> of 1.0 mol dm<sup>-3</sup> sulphuric acid and make up to 1.0 dm<sup>3</sup> with distilled water)
- Access to 0.10 mol dm<sup>-3</sup> iron(III) chloride solution (make by dissolving 16.2 g of anhydrous iron(III) chloride (**Harmful**) in distilled water and make up to 1.0 dm<sup>3</sup>)
- Access to starch solution
- Access to approx. 0.50 mol dm<sup>-3</sup> potassium iodide solution (make by dissolving 83 g of potassium iodide in 1.0 dm<sup>3</sup> distilled water)

## Safety

The main points are included on the Student Sheet but it is the teacher's responsibility to ensure that a full risk assessment is carried out prior to the practical session. As there are some hazards associated with the solutions used, safety issues should be stressed, and use of eye protection made mandatory. MSDS sheets should be consulted so that the correct action can be taken in event of a spillage and/or accident. Materials that are dangerous to the environment should be disposed of according to local regulations.