B14 · Investigation: plan and predict

Completed notes

Goals: Write a testable question with the independent, dependent and controlled variables. Draw a labelled cell design and predict the anode, the equation and the voltage from the table. Plan safe handling and disposal and plan a battery with cells in a row.

Slide 1 of 8 · A testable question and its variables

A testable question links one thing we **change** to one thing we **measure**. Here we change the metal in one half-cell and measure the cell voltage.

The independent variable is the one we **choose to change**: the metal. The dependent variable is the one we measure: the **voltage**.

Controlled variables stay the **same** for a fair test. The table values are for 1.0 mol/L solutions, so we control the **concentration**.

We also control the temperature, the electrode **size** and the salt bridge. If two things changed, we could not tell which one caused the voltage change.

Example. We want to test the reduction table with real cells. Each metal is joined to a copper half-cell in copper(II) sulfate. Write a testable question. Name the independent variable, the dependent variable and two controlled variables.

1. Question: How does the metal in one half-cell change the voltage of the cell?

2. Independent variable: the metal. Dependent variable: the voltage in volts.

3. Controlled variables: the concentration (1.0 mol/L) and the temperature. The electrode size and the salt bridge also stay the same.

Try it. A student asks whether the concentration of the copper(II) sulfate solution changes the voltage of a zinc and copper cell. Name the independent variable, the dependent variable and two controlled variables.

1. Independent variable: the concentration of the copper(II) sulfate solution.

2. Dependent variable: the voltage of the cell.

3. Controlled variables: any two of the metals (zinc and copper), the temperature, the electrode size and the salt bridge.

Slide 2 of 8 · A design with a labelled diagram

Half-cell 1 is a strip of metal X in a 1.0 mol/L solution of its **ions**. Half-cell 2 is a copper strip in 1.0 mol/L **copper(II) sulfate**.

A salt bridge joins the solutions. It lets **ions** move so that charge stays balanced and the circuit is complete. Wires join the strips through a **voltmeter**.

Predict first, then label. The half-cell with the **lower** reduction value is the anode, because that metal is oxidized. The **higher** value is the cathode.

Electrons flow in the wire from the **anode** to the **cathode**. A good diagram also labels the solutions with their concentrations.

Example. We will join each metal half-cell to a copper half-cell. Draw the design for metal X in $X^{2+}(aq)$ joined to copper in $Cu^{2+}(aq)$. Label the parts. Assume X has the lower reduction value, and show the direction of electron flow.

1. Half-cell 1: X(s) in 1.0 mol/L $X^{2+}(aq)$. Half-cell 2: Cu(s) in 1.0 mol/L $Cu^{2+}(aq)$.

2. A salt bridge joins the two solutions. Wires join the strips through a voltmeter.

3. X is the anode (oxidation). Copper is the cathode (reduction).

4. Electrons flow in the wire from X to copper.

Try it. Draw a design to test a silver half-cell against a copper half-cell, both 1.0 mol/L. Given $E°(Cu^{2+}/Cu) = +0.34$ V and $E°(Ag^{+}/Ag) = +0.80$ V, label the anode, the cathode and the electron flow.

1. Half-cells: Cu(s) in 1.0 mol/L $Cu^{2+}(aq)$ and Ag(s) in 1.0 mol/L $Ag^{+}(aq)$, with a salt bridge, wires and a voltmeter.

2. Copper has the lower value (+0.34 V), so copper is the anode.

3. Silver has the higher value (+0.80 V), so silver is the cathode.

4. Electrons flow in the wire from copper to silver.

Slide 3 of 8 · Predict before measuring

To predict, compare the two table values. The metal with the **lower** value is the anode and is oxidized. The metal with the higher value is the **cathode**.

For zinc and copper, $E°(Zn^{2+}/Zn) = -0.76$ V and $E°(Cu^{2+}/Cu) = +0.34$ V. The cell reaction is $Zn + Cu^{2+} \rightarrow \mathbf{Zn^{2+} + Cu}$. Charge is +2 on each side.

The voltage is $E°cell = E°cathode - E°anode = 0.34 - (-0.76) = \mathbf{+1.10\ V}$. A flipped sign would give a **negative** cell voltage, which is wrong.

Copper is not always the cathode. Silver (+0.80 V) is higher than copper, so silver is the **cathode** and copper is the **anode**. Compare the two values **every time**.

Example. Predict the anode and E°cell when each metal is joined to a 1.0 mol/L copper half-cell. Given $E°(Zn^{2+}/Zn) = -0.76$ V, $E°(Fe^{2+}/Fe) = -0.45$ V, $E°(Ni^{2+}/Ni) = -0.26$ V, $E°(Ag^{+}/Ag) = +0.80$ V and $E°(Cu^{2+}/Cu) = +0.34$ V, do zinc, iron, nickel and silver.

1. Zinc: lower than copper, so zinc is the anode. $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

2. $E°cell = 0.34 - (-0.76) = +1.10$ V

3. Iron: iron is the anode. $E°cell = 0.34 - (-0.45) = +0.79$ V

4. Nickel: nickel is the anode. $E°cell = 0.34 - (-0.26) = +0.60$ V

5. Silver: higher than copper, so silver is the cathode and copper is the anode.

6. $E°cell = 0.80 - 0.34 = +0.46$ V

Try it. Given $E°(Fe^{2+}/Fe) = -0.45$ V and $E°(Ag^{+}/Ag) = +0.80$ V. An iron half-cell is joined to a silver half-cell, both 1.0 mol/L. Predict the anode, the cathode and E°cell.

1. Iron has the lower value, so iron is the anode. Silver has the higher value, so silver is the cathode.

2. $E°cell = 0.80 - (-0.45) = +1.25$ V

Slide 4 of 8 · Safe handling, labels and disposal

Read the label and the **safety data sheet** first. They tell the hazards and the first aid. They come from **WHMIS**, the Workplace Hazardous Materials Information System.

Plan protection: **goggles**, gloves and an apron. Silver nitrate stains **skin**. Copper(II) compounds harm life in water.

Store solutions in **closed, labelled** containers, so nobody mixes them up or spills them.

Collect solutions with metal ions in a **waste container**. Do not pour them down the **drain**, because metal ions can harm water life. Follow the school disposal plan.

Example. A plan uses copper(II) sulfate and silver nitrate solutions. Name one hazard of each with a matching precaution. Then say how to handle the waste.

1. Silver nitrate stains skin, so wear gloves. Read the label and the safety data sheet first.

2. Copper(II) sulfate harms life in water, so it must not go down the drain.

3. Wear goggles and an apron while working.

4. Collect all metal-ion waste in a waste container and follow the school disposal plan.

Try it. A plan uses nickel(II) sulfate solution and nickel strips. Name one hazard and a matching precaution, one storage step and one disposal step. Which item can be reused?

1. Nickel(II) compounds are harmful, so wear goggles, gloves and an apron.

2. Store the solution in a closed, labelled container.

3. Collect the waste in a waste container. Do not pour it down the drain.

4. The nickel strips can be rinsed, dried and reused.

Slide 5 of 8 · Method: how to plan and predict an investigation

Step 1. Write a **testable** question. It links one thing you change to one thing you measure.

Step 2. Name the variables. **Manipulated**: what you change. **Responding**: what you measure. Controlled: what stays the same.

Step 3. Draw a labelled **design**. Show the solutions with concentrations, the salt bridge, the wires and the meter.

Step 4. **Predict** with the table. Find the anode, the cell reaction and E°cell before you measure.

Step 5. List **safety** and disposal. Name a hazard, a precaution and how the waste is collected.

Example. Use the method on testing the reduction table by joining each metal to a copper half-cell. Plan the investigation and predict the result for zinc.

1. Step 1: how does the cell voltage change with the metal?

2. Step 2: manipulated is the metal. Responding is the voltage. Controlled are 1.0 mol/L solutions, temperature, electrode size and salt bridge.

3. Step 3: two labelled half-cells, a salt bridge, wires and a voltmeter.

4. Step 4: Zn is the anode. $E°cell = 0.34 - (-0.76) = +1.10$ V.

5. Step 5: goggles, gloves and an apron. Collect metal-ion waste.

Try it. Use the method on joining a nickel half-cell to a silver half-cell, both 1.0 mol/L. $E°(Ni^{2+}/Ni) = -0.26$ V and $E°(Ag^{+}/Ag) = +0.80$ V. Predict the anode and the voltage, and name one hazard.

1. Ni is lower, so Ni is the anode and Ag is the cathode.

2. $E°cell = 0.80 - (-0.26) = +1.06$ V.

3. Hazard: silver nitrate stains skin, so wear gloves and collect the waste.

Slide 6 of 8 · Your turn: nickel and silver

Compare the two table values. The **lower** value belongs to the metal that is oxidized, so that metal is the **anode**. Nickel (−0.26 V) is lower than silver (+0.80 V).

Balance the electrons first. $Ni \rightarrow Ni^{2+} + 2e^-$ gives 2 electrons, and each Ag^+ takes **1**, so double the silver half: $2Ag^+ + 2e^- \rightarrow$ **2Ag**.

Add the halves: $Ni + 2Ag^+ \rightarrow Ni^{2+} + 2Ag$. Check that the charge is **+2** on each side. Do not **multiply** E° by the coefficient.

$E°cell = 0.80 - (-0.26) =$ **+1.06 V**. Give the voltage with its sign and its **unit**. Silver nitrate stains skin, so wear **gloves**.

Example. Given a plan to join a nickel half-cell, $E°(Ni^{2+}/Ni) = -0.26$ V, to a silver half-cell, $E°(Ag^+/Ag) = +0.80$ V, both 1.0 mol/L. Predict the anode, the cell reaction and the voltage. Name one hazard.

1. The lower value is nickel, so nickel is the anode.

2. Oxidation: $Ni \rightarrow Ni^{2+} + 2e^-$

3. Reduction: $2Ag^+ + 2e^- \rightarrow 2Ag$

4. Cell reaction: $Ni + 2Ag^+ \rightarrow Ni^{2+} + 2Ag$. Check: Ni 1, Ag 2, charge +2 on each side.

5. $E°cell = 0.80 - (-0.26) = +1.06$ V

6. Hazard: silver nitrate stains skin, so wear gloves.

Try it. Given a plan to join a zinc half-cell, $E°(Zn^{2+}/Zn) = -0.76$ V, to a silver half-cell, $E°(Ag^+/Ag) = +0.80$ V, both 1.0 mol/L. Predict the anode, the cell reaction and the voltage. Name one disposal step.

1. The lower value is zinc, so zinc is the anode.

2. Oxidation: $Zn \rightarrow Zn^{2+} + 2e^-$. Reduction: $2Ag^+ + 2e^- \rightarrow 2Ag$

3. Cell reaction: $Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$. Check: Zn 1, Ag 2, charge +2 on each side.

4. $E°cell = 0.80 - (-0.76) = +1.56$ V

5. Disposal: collect the metal-ion waste in a container and do not pour it down the drain.

Slide 7 of 8 · Plan a battery

Cells in a row **add** their voltages. One zinc and copper cell gives **1.10 V**, and the device needs at least 3.0 V.

Divide: 3.0 ÷ 1.10 = 2.73. You cannot use part of a cell, so **round up** to three cells. Two cells give only 2.20 V, which is not enough.

Three cells give 3 × 1.10 = **3.30 V**, which is more than 3.0 V. The table gives the voltage at the start only. As reactants are used up, the voltage **falls**.

So the plan needs a **test**: record the voltage over time while the device runs. A salt bridge adds **resistance**, and magnesium gives more voltage but reacts with water.

Example. A small device needs at least 3.0 V. One zinc and copper cell gives E°cell = +1.10 V. How many cells in a row are needed? What must the plan check?

1. Number of cells = 3.0 V ÷ 1.10 V per cell = 2.73, so round up to 3 cells.

2. Check: 3 × 1.10 = 3.30 V is enough. 2 × 1.10 = 2.20 V is not enough.

3. The plan must test that the voltage stays at or above 3.0 V while the device runs, because the voltage falls as the cell is used.

Try it. A device needs at least 4.0 V. One nickel and copper cell gives E°cell = +0.60 V. How many cells in a row are needed? Why might the real battery give less voltage than the table predicts?

1. Number of cells = 4.0 V ÷ 0.60 V per cell = 6.67, so round up to 7 cells.

2. Check: 7 × 0.60 = 4.20 V is enough. 6 × 0.60 = 3.60 V is not enough.

3. The voltage falls as reactants are used up, and a salt bridge adds resistance.

Slide 8 of 8 · What a plan cannot show

A prediction uses the **table** and tells what we expect. An observation comes from a **measurement** and tells what really happens.

A plan only describes how we will get the measurement. A plan is **neither** a prediction nor an observation, and it is not **data**.

A prediction becomes a test of an idea only when we **compare** it with an observation. A common wrong choice treats a predicted value as an **observation**.

Table values apply only at standard conditions: 1.0 mol/L solutions, **25.00 °C** and 101.325 kPa for gases. Real cells differ a little, so measurements differ a **little** from predictions.

Example. A report lists 1.10 V for a zinc and copper cell under the heading Measured voltage. The value came from the reduction table. Is it an observation or a prediction? How can you tell? Why might a real measurement differ?

1. It is a prediction, because it was calculated from the table: $0.34 - (-0.76) = 1.10$ V.

2. It did not come from a voltmeter, so it is not an observation.

3. A real cell can differ because of salt bridge resistance, oxide on the strips, or solutions that are not exactly 1.0 mol/L at 25.00 °C.

Try it. A given sample data set lists an iron and copper cell with a table value of 0.79 V and a voltmeter reading of 0.76 V. These readings were not measured by us. Which number is the prediction and which is the observation? Is the plan itself either one?

1. The table value, 0.79 V, is the prediction. The voltmeter reading, 0.76 V, is the observation.

2. The plan is neither one. It only describes how the measurement would be made.

B14 · Worked practice

Investigation: plan and predict | Questions 1-2 of 7

1. Given the question "How does the metal in one half-cell change the cell voltage?" Name the independent variable, the dependent variable and two controlled variables.

Independent: the metal. Dependent: the voltage. Controlled: any two of the concentration (1.0 mol/L), the temperature, the electrode size and the salt bridge.

2. Given a plan to join a zinc half-cell to a copper half-cell with a salt bridge and a voltmeter. List the parts that a labelled diagram must show.

Two half-cells with solutions, concentrations and strips, the salt bridge, the wires and the voltmeter. Label the anode, the cathode and the electron flow.

B14 · Worked practice

Investigation: plan and predict | Questions 3-4 of 7

3. Given the table values $E°(Fe^{2+}/Fe) = -0.45$ V and $E°(Cu^{2+}/Cu) = +0.34$ V. Predict the anode, the cell reaction and the voltage.

Use Method: how to plan and predict an investigation. Iron is the anode. $Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$. $E°cell = 0.34 - (-0.45) = +0.79$ V.

4. Given a plan that uses silver nitrate solution. Name two hazards or precautions and one disposal step.

Silver nitrate stains skin, so wear gloves, goggles and an apron. Read the label and safety data sheet. Collect the waste in a container and do not pour it down the drain.

B14 · Worked practice

Investigation: plan and predict | Questions 5-6 of 7

5. Given a nickel half-cell and a copper half-cell, both 1.0 mol/L. Predict the anode and the voltage (Data Booklet table values: $E°(Ni^{2+}/Ni) = -0.26$ V; $E°(Cu^{2+}/Cu) = +0.34$ V).

Nickel is the anode. $E°_{cell} = 0.34 - (-0.26) = +0.60$ V.

6. Given a device that needs at least 4.0 V and a cell of 1.10 V. Find the number of cells in a row, and give one thing that the plan should test.

4.0 ÷ 1.10 = 3.6, so four cells (4.40 V). The plan should test that the voltage stays above 4.0 V while the device runs, because the voltage falls as the cell is used.

B14 · Worked practice

Investigation: plan and predict | Questions 7-7 of 7

7. Challenge: Given a device that needs at least 6.0 V. Two cell designs are possible. Data Booklet table values: $E°(Zn^{2+}/Zn) = -0.76$ V, $E°(Mg^{2+}/Mg) = -2.37$ V and $E°(Cu^{2+}/Cu) = +0.34$ V. Design 1 joins zinc and copper half-cells. Design 2 joins magnesium and copper half-cells. Predict the voltage of one cell of each design. Then find the number of cells in a row that each design needs.

Design 1: copper is the cathode and zinc is the anode. $E°cell = +0.34 - (-0.76) = +1.10$ V. Cells needed: $6.0 \div 1.10 = 5.45$, so round up to 6 cells (6.60 V). Design 2: copper is the cathode and magnesium is the anode. $E°cell = +0.34 - (-2.37) = +2.71$ V. Cells needed: $6.0 \div 2.71 = 2.21$, so round up to 3 cells (8.13 V). Design 2 needs fewer cells. Always round up: one cell too few gives less than 6.0 V.