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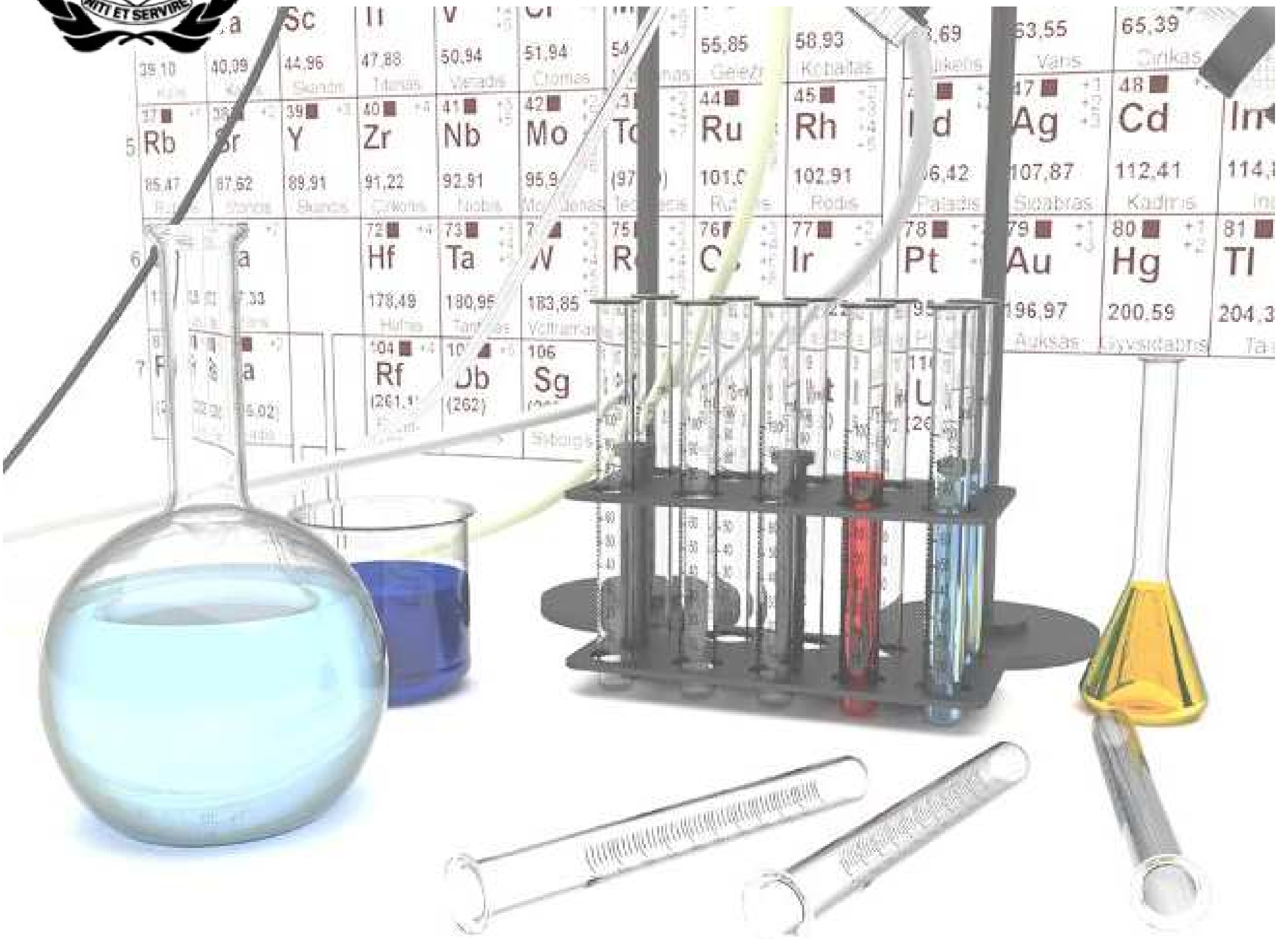

METALS and ELECTROCHEMISTRY

1

Including:

Section 1 : METALS and REACTIVITY

Section 2 : REDOX

Section 3 : GALVANIC CELLS

Section 4 : ELECTROLYSIS

Student booklet

Name: ___________________________

Teacher: ___________________________

Metal Extraction Chemistry

Standard electrode potentials at 25°C[a]

oxidising strength increases ↑ ; reducing strength increases ↓

Oxidised form + ne^-	→	Reduced form	$E^\ominus$ (V)
$F_2 + 2e^-$	→	$2F^-$	+2.87
$H_2O_2 + 2H^+ + 2e^-$	→	$2H_2O$	+1.78
$Au^+ + e^-$	→	Au	+1.69
$MnO_4^- + 8H^+ + 5e^-$	→	$Mn^{2+} + 4H_2O$	+1.51
$PbO_2 + 4H^+ + 2e^-$	→	$Pb^{2+} + 2H_2O$	+1.46
$Cl_2 + 2e^-$	→	$2Cl^-$	+1.36
$O_3 + H_2O + 2e^-$	→	$O_2 + 2OH^-$	+1.24
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	→	$2Cr^{3+} + 7H_2O$	+1.23
$O_2 + 4H^+ + 4e^-$	→	$2H_2O$	+1.23
$MnO_2 + 4H^+ + 2e^-$	→	$Mn^{2+} + 2H_2O$	+1.22
$Pt^{2+} + 2e^-$	→	Pt	+1.18
$Ag_2O + 2H^+ + 2e^-$	→	$2Ag + H_2O$	+1.17
$Br_2 + 2e^-$	→	$2Br^-$	+1.09
$NO_3^- + 4H^+ + 3e^-$	→	$NO + 2H_2O$	+0.96
$NO_3^- + 3H^+ + 2e^-$	→	$HNO_2 + H_2O$	+0.93
$2Hg^{2+} + 2e^-$	→	Hg_2^{2+}	+0.92
$Hg^{2+} + 2e^-$	→	Hg	+0.85
$NO_3^- + 2H^+ + e^-$	→	$NO_2 + H_2O$	+0.80
$Ag^+ + e^-$	→	Ag	+0.80
$Fe^{3+} + e^-$	→	Fe^{2+}	+0.77
$O_2 + 2H^+ + 2e^-$	→	H_2O_2	+0.70
$MnO_4^- + 2H_2O + 3e^-$	→	$MnO_2 + 4OH^-$	+0.60
$I_2 + 2e^-$	→	$2I^-$	+0.54
$O_2 + 2H_2O + 4e^-$	→	$4OH^-$	+0.40
$Ag_2O + H_2O + 2e^-$	→	$2Ag + 2OH^-$	+0.34
$Cu^{2+} + 2e^-$	→	Cu	+0.34
$SO_4^{2-} + 4H^+ + 2e^-$	→	$H_2SO_3 + H_2O$	+0.17
$Sn^{4+} + 2e^-$	→	Sn^{2+}	+0.15
$S + 2H^+ + 2e^-$	→	H_2S	+0.14
$S_4O_6^{2-} + 2e^-$	→	$2S_2O_3^{2-}$	+0.08
$2H^+ + 2e^-$	→	H_2	0.00
$Pb^{2+} + 2e^-$	→	Pb	−0.13
$Sn^{2+} + 2e^-$	→	Sn	−0.14
$Ni^{2+} + 2e^-$	→	Ni	−0.26
$Co^{2+} + 2e^-$	→	Co	−0.28
$PbSO_4 + 2e^-$	→	$Pb + SO_4^{2-}$	−0.36
$Cd^{2+} + 2e^-$	→	Cd	−0.40
$Fe^{2+} + 2e^-$	→	Fe	−0.45
$2CO_2 + 2H^+ + 2e^-$	→	$H_2C_2O_4$	−0.49
$Zn^{2+} + 2e^-$	→	Zn	−0.76
$2H_2O + 2e^-$	→	$H_2 + 2OH^-$	−0.83
$Al^{3+} + 3e^-$	→	Al	−1.66
$Mg^{2+} + 2e^-$	→	Mg	−2.37
$Na^+ + e^-$	→	Na	−2.71
$Ca^{2+} + 2e^-$	→	Ca	−2.87
$Ba^{2+} + 2e^-$	→	Ba	−2.91
$K^+ + e^-$	→	K	−2.93
$Li^+ + e^-$	→	Li	−3.04

Worksheet 1.1

Metal Extraction Chemistry

METALS

Please answer the following questions in point form.

a) What is a Metal?

b) Where are metals found on the Periodic table?

c) As a group metals have specific chemical properties related to their atomic structure. Explain the unique features of metals in terms of their…
   a. Electron configuration
   b. Ionisation Energy
   c. Electronegativity
   d. Density

d) Metals have been used for thousands of years. Complete the table below which relates a types of metal with its use and the reason the metal is used.

Metal	Symbol	Use	Property which make it useful
Gold			
Silver			
Iron			
Zinc			
Copper			
Mercury			
Lead			
Aluminium			
Tin			
Tungsten			

e) When were various metals discovered? Place the metals named in the table above in one of the columns of the table below

Discovered More than 2000 yrs ago	Discovered between 200 and 2000 yrs ago	Discovered in last 200 yrs

f) Now the IMPORTANT question? Look closely at the above three groups of metals. What similarities can you see within each group and what differences are there between groups? Explain why the metals were found at these different times!

g) What are alloys and why are alloys sometimes more useful than pure metals?

Metal Extraction Chemistry

EXPERIMENT 1.1

COMPARING THE REACTIVITY OF DIFFERENT METALS

INTRODUCTION

Metals display a wide range of reactivity with many other substances. Relatively unreactive metals such as gold and silver may be found in their metallic state in nature. Reactive metals such as aluminium and sodium, although quite plentiful, are always found in compounds. In this experiment, you will determine the relative reactivity of a number of different metals and arrange them in order of decreasing reactivity. Here we shall use displacement reactions.

AIM

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EQUIPMENT

small pieces of copper, lead, iron, magnesium, zinc

20 mL of each of the following 0.1 mol/L solutions

- copper(II) sulfate (CuSO4)
- iron(II) sulfate (FeSO4)
- lead(II) nitrate (Pb(NO3)2)
- zinc sulfate (ZnSO4)
- magnesium sulfate (MgSO4)
- silver nitrate (AgNO3)

25 test tubes sandpaper test tube racks

SAFETY

Wear safety glasses. Lead nitrate is toxic so avoid contact with skin. Wash with plenty of water if contact occurs. Silver nitrate will stain skin so avoid contact. Lead and silver compounds should not be disposed of down the drain.

PROCEDURE

1 Place five test tubes in a rack and add approximately 3 mL of each of the solutions to separate test tubes. (There is no need to test the metal with a solution of the same ion.)

2 Add to each of the five test tubes a small piece of one of the metals, which has been first cleaned with sandpaper. (Do not add the same metal to its own solution; for example, do not place magnesium in a magnesium sulfate solution.)

3 Observe and record results. If no change is visible within 5 minutes, write down 'no reaction'.

4 Repeat the above procedure with all other metals. Wash the pieces of metal carefully after each experiment if they are being re-used.

Metal Extraction Chemistry

RESULTS

Solution/metal	Copper	Iron	Lead	Zinc	Magnesium
Copper sulfate					
Iron sulfate					
Lead nitrate					
Zinc sulfate					
Magnesium sulfate					
Silver nitrate					

QUESTIONS

1 a Which of the metals gave:

i five reactions? ________________

ii four reactions? ________________

iii three reactions? _______________

iv two reactions? _______________

v one reaction? ________________

b Use the above to list the five metals in order from most reactive to least reactive.

____

2 In all the reactions, one metal displaces a less active one from solution.

Write an equation for the reaction between:

a zinc and copper sulphate ____

b iron and lead nitrate ____

3 Consider the reactivity of silver relative to the other metals. Where would you place silver in the activity list? ____

CONCLUSION (trends/patterns summarised, conclusions made, validity of conclusion ie errors)

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Source: Debra Smith, *Conquering Chemistry Preliminary Course Blackline Masters*, Worksheet 7, McGraw-Hill Australia, 2003.

WORKSHEET 1.2

THE ACTIVITY SERIES OF METALS

Definition:

The Activity series:

Reactivity	Metal/Element	Symbol
High Activity Most Reactive		
Decreasing activity/ reactivity		
Low Activity Least Reactive		

Metal Extraction Chemistry

WORKSHEET 1.3

ACTIVITY OF METALS

1 By referring to the Activity Series, select four metals that will displace:

a Tin from a solution of Tin nitrate ______________________________

b copper from a solution of copper sulphate __________________________________

2 Write balanced net ionic equations for the following:

a Zinc is placed in a copper sulfate solution.

____

b Lead is added to a solution containing Ag^+ ions.

____

c Iron is placed in a solution containing Na^+ and Pb^{2+} ions

____

3 Four metals are designated W, X, Y, Z. Solid samples of W, X and Y are placed in solutions of a salt of Z. W and X remain unchanged, but Y becomes coated with a solid which on testing is found to be Z. X does not react with dilute hydrochloric acid, but W dissolves, giving hydrogen gas.

a Draw up a table of results.

b Arrange these metals in order from most reactive to least reactive.

____

c List the metals in order of reducing strength, beginning with the strongest reductant.

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d If only one of the metals is found free in nature, which is it likely to be?

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Source: Debra Smith, *Conquering Chemistry HSC Course Blackline Masters*, Worksheet 10, McGraw-Hill Australia, 2003.

WORKSHEET 1.4

OXIDATION and REDUCTION

1. Define the following terms:
   a) Redox:

   b) Oxidation:

   c) Reduction:

   d) Oxidant:

   e) Reductant:

   f) Half Reaction:

   g) Oxidation Number:

   h) Redefine Oxidation and Reduction (in terms of changes in oxidation number):

2. Make a list of the rules for assigning Oxidation Numbers:

Metal Extraction Chemistry

WORKSHEET 1.5

OXIDATION AND REDUCTION

1 Classify each of the following statements as true (T) or false (F). For those statements that are false rewrite the statement so it is correct.

a Reduction is the gain of electrons by a substance. ___________

b The oxidation number of an uncombined element is always zero. ___________

c When copper loses two electrons to form Cu^{2+} it is reduced. __________

d An oxidation reaction is always accompanied by a reduction reaction. _________

e An oxidising agent or oxidant is the substance that is oxidised. __________

f For positive monatomic ions the oxidation state is always +1. __________

g The oxidation number of manganese in MnO_2 is +1. _________

h For any redox reaction the no. of electrons lost must be equal to the no. of electrons gained. ________

i In the reaction $CuO_{(s)} + H_{2(g)} \rightarrow Cu_{(s)} + H_2O_{(l)}$, Cu goes from an oxidation state of +1 to 0. ________

j For any neutral compound the sum of the oxidation numbers within the molecule must equal zero. ________

3 Give the oxidation number of each element in the following:

a potassium bromide ____

b magnesium ____

c aluminium oxide ____

d iron(II) chloride ____

e iodine ____

f iron(III) chloride ____

4 Identify the species that is oxidised and the one that is reduced in each of the following reactions, and name the oxidant and reductant:

a $Cl_2 + 2HBr \rightarrow 2HCl + Br_2$

____

____

b $I_2O_5 + 3CO \rightarrow I_2 + 3CO_2$

____

____

c $6Mn^{3+} + I^- + 6OH^- \rightarrow 6Mn^{2+} + IO_3^- + 3H_2O$

____

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Metal Extraction Chemistry

5. Balance the following redox equations:

a $HNO_3 + HI \rightarrow NO + I_2 + H_2O$

b $SO_2 + H_2S \rightarrow H_2O + S$

c $V_2O_5 + HCl \rightarrow VOCl + H_2O + Cl_2$

7 Consider the following reactions of unknown cations M^+ and N^{2+} with halogens in aqueous solution where M^+ is oxidised to M^{2+} by Cl_2 and Br_2, but not I_2. N^{2+} is oxidised to N^{4+} by Cl_2, but not by Br_2 or I_2. The reactions of F_2 were not tested since F_2 oxidises water. Using this information, arrange the halogens in order of increasing oxidising strength.

8 Consider the following substances X_2, Y_2 and Z_2. H_2S is converted to S by X_2, Y_2 and Z_2; Fe^{2+} is converted to Fe^{3+} by X_2 and Y_2 but not Z_2; M^+ is converted to M^{3+} by X_2, but not by Y_2 or Z_2.

On the basis of this data, which of the following reactions are likely to proceed?

A $X_2 + 2Y^- \rightarrow 2X^- + Y_2$

B $Y_2 + 2Z^- \rightarrow 2Y^- + Z_2$

C $Z_2 + 2X^- \rightarrow 2Z^- + X_2$

Source: Debra Smith, *Conquering Chemistry Preliminary Course Blackline Masters*, Worksheet 7, McGraw-Hill Australia, 2003.

WORKSHEET 1.6

ELECTRICITY FROM REDOX REACTION

(Galvanic Cells / Batteries)

1. Why can redox reactions "make electricity" (generate an electric current)

____

____

____

2. To understand Galvanic cells, the following definitions are important:

(i) Galvanic Cells ______________________________

____

(ii) Electrode ______________________________

____

(iii)Electrolyte ______________________________

____

(iv)Anode ______________________________

____

(v) Cathode ______________________________

____

(vi)Half-cell ______________________________

____

(vii) Salt bridge ______________________________

____

(viii) EMF (cell voltage) ______________________________

____

____

(ix) Standard Electrode Potential ______________________________

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In your owns if possible write down: Firstly why galvanic cells produce an electric current; and secondly, what determines the voltage they can produce.

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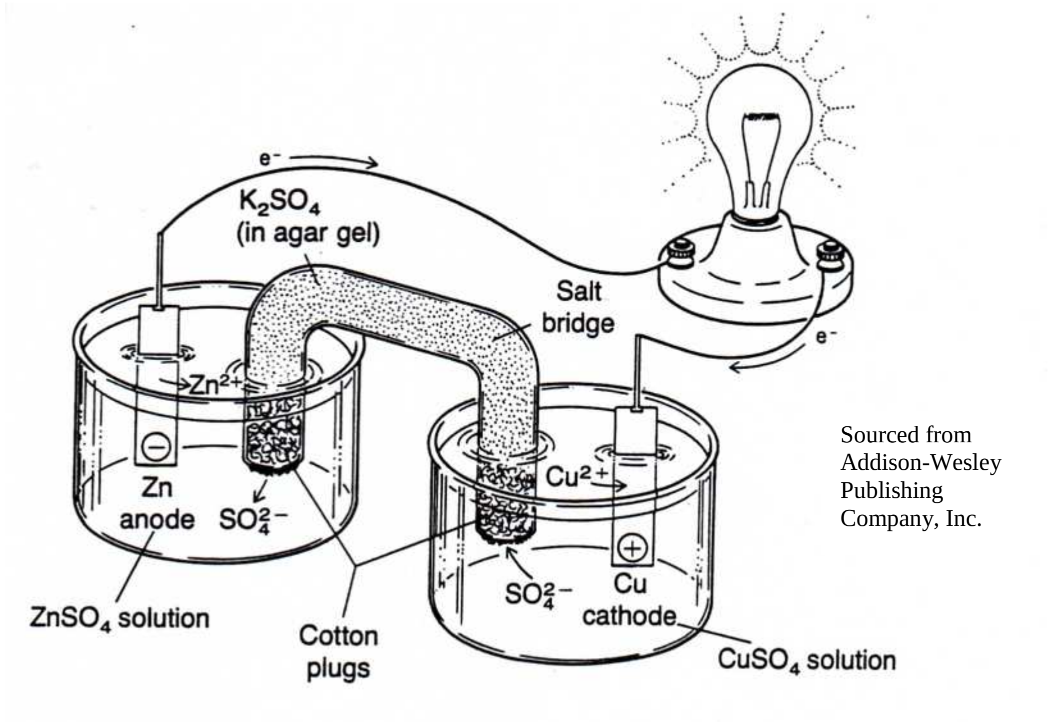
WORKSHEET 1.7

GALVANIC CELLS

Galvanic/Voltaic/Electrochemical cell:



Sourced from Addison-Wesley Publishing Company, Inc.

1. Determine the location where each process listed below occurs in the diagram above. Write the letter of that process on the diagram in large bold letters.
    A. Electrons pass trough the external circuit to the copper strip
    B. Positive and negative ions move through the aqueous solution to maintain electrical neutrality
    C. Electrons are passed to copper ions, and reduction takes place
    D. Electrons are released by oxidation occuring

2. Write the half-reactions and overall cell reaction for this Galvanic / Voltaic / Electrochemical cell:

    Oxidation half-reaction: ____

    Reduction half-reaction: ____

    Overall Cell reaction: ____

EXPERIMENT 1.2

CONSTRUCTING GALVANIC CELLS

INTRODUCTION

For a galvanic cell to work the components of the redox reaction must be separated so that the electron transfer fl ows through an external circuit. The circuit is completed by a salt bridge containing a suitable electrolyte. In this experiment the salt bridge consists of a strip of fi lter paper soaked in potassium nitrate solution.

AIM

To construct different galvanic cells and measure the differences in voltage when different combinations of metals are used.

EQUIPMENT

- 1 mol/L zinc nitrate ($Zn(NO_3)_2$)
- 1 mol/L copper nitrate ($Cu(NO_3)_2$)
- 1 mol/L lead nitrate ($Pb(NO_3)_2$)
- 1 mol/L iron(II) sulfate ($FeSO_4$)
- 1 mol/L magnesium sulfate ($MgSO_4$)
- saturated potassium nitrate (KNO_3) solution
- 5 × 100 mL beakers
- 1 × 250 mL beaker
- stirring rod
- voltmeter
- connecting wires
- fine sandpaper or steel wool
- strips of zinc, copper, iron, lead and magnesium metal
- strips of filter paper about 1 cm wide and 8 cm long

SAFETY

Wear safety glasses. Lead salts are poisonous. If contact with skin occurs wash with plenty of water. Do not dispose of lead solutions down the sink.

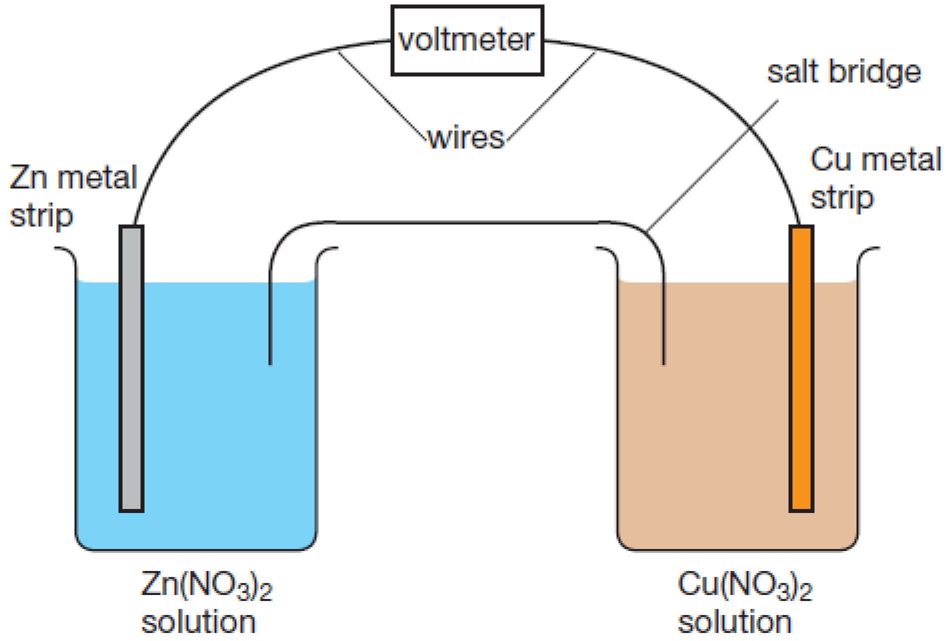
PROCEDURE

1 Clean the metal strips with sandpaper or steel wool.

2 Place the solutions of zinc nitrate, copper nitrate, lead nitrate, iron sulfate and magnesium sulfate into separate 100 mL beakers (about half-full). Add the same metal to the beaker containing the metal ion solution; for example, zinc metal goes in zinc nitrate solution.

3 These are your half-cells. Put strips of filter paper in a 250 mL beaker, add saturated potassium nitrate solution and leave to soak. These are your salt bridges.

4 Connect the zinc half-cell to the copper half-cell with the salt bridge.



5 Use the connecting wires to connect the zinc metal and copper metal to the voltmeter as shown in the diagram above. You may need to swap the wires on the voltmeter to get a reading. Note which electrode (metal) is positive and the magnitude of the voltage; record these observations in the results table.

6 Remove the salt bridge and note what happens to the voltage.

7 Repeat with the other half-cells, trying different combinations. Use a new salt bridge each time to avoid contamination. Note the voltage when the salt bridge is removed for the first three cells only, and assume the effect is the same for all cells.

RESULTS

Beaker 1/beaker 2	Polarity	Voltage	Anode reaction	Cathode reaction
Zinc/copper	**- / +**		**Zn** $\rightarrow$ Zn^{+2}	**Cu^{+2}** $\rightarrow$ **Cu**
Zinc/iron				
Iron/copper				
Lead/magnesium				

QUESTIONS

1 Complete the table above, giving the anode and cathode reactions for each combination.

2 Write the overall reaction for five of the above combinations.

3 a Why is a salt bridge necessary?

b Describe the flow of ions through the salt bridge—use the terms 'anions' and 'cations'.

4. Compare the voltage of the cells with the positions of the metals in the activity series. Is there a relationship?

5 Use a table of standard reduction potentials to calculate the E° value for each cell.

Compare that value with the value obtained in the experiment. Suggest reasons for any differences.

Source: Debra Smith, Conquering Chemistry HSC Course Blackline Masters, Worksheet 11, McGraw-Hill Australia, 2003.

WORKSHEET 1.8

GALVANIC CELLS

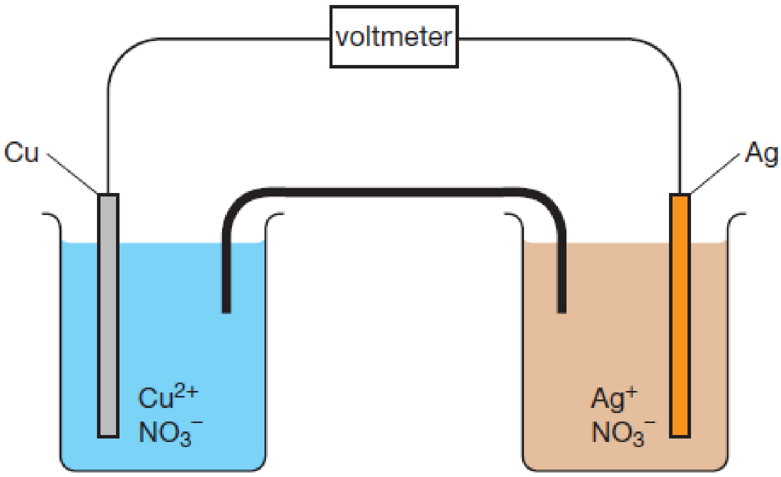
1 Check your understanding by completing the following:

A _____________ cell is a device in which a chemical reaction occurs in such a way that it generates ________________ . The electrode at which oxidation occurs is the _______________ and the electrode at which reduction occurs is the ________________ .

In a galvanic cell the ____________ reactions occur at different locations. The solution in a half cell is called an __________________ . The half-cell solutions are connected by a _____________ bridge, which permits the passage of ______________ between them. The positive ions or _________________ flow towards the ________________ while the ________________ ions or anions flow towards the ____________ .

The conductors of a cell that get connected to the external circuit are called __________________ . Reduction occurs at the ________________ , the positively charged electrode while _________________ occurs at the _________________ , the __________________ charged electrode. Electrons flow via an external circuit from the ________________ to the __________________ . The cell voltage is the difference between the ________________ potential of the half-cells.

2 On the diagram below, given that Ag is the positive electrode, label the cathode, the anode, the direction of electron flow and the salt bridge.



3 Given the following two half-cell reactions, write the balanced overall reaction.

$Cu_{(s)} \rightarrow Cu^{2+}{}_{(aq)} + 2e^-$ $\qquad$ $Ag^{+}{}_{(aq)} + e^- \rightarrow Ag_{(s)}$

____

4 Given the overall cell reaction, write the two half-cell reactions and identify the oxidation and reduction reactions.

$3Pb^{2+}{}_{(aq)} + 2Cr_{(s)} \rightarrow 3Pb_{(s)} + 2Cr^{3+}{}_{(aq)}$

5 From the following shorthand representation, identify the anode and cathode, and write the oxidation and reduction half-reactions given that Pb is positive.

$Fe_{(s)}$ | $Fe^{+2}{}_{(aq)}$ || $Pb^{+2}{}_{(aq)}$ | $Pb_{(s)}$

6 Write the half-reactions, the overall reaction and the shorthand representation for the following electrochemical cell.

One electrode consists of a piece of silver dipping into a silver nitrate solution. It is connected by a salt bridge to another electrode, which consists of a piece of platinum dipping into a chloride solution with chlorine gas bubbling through the solution over the inert platinum electrode. The platinum is the positive electrode.

7 Draw a diagram of an electrochemical cell with the following overall reaction:

$2Al_{(s)} + 3Cu^{2+}{}_{(aq)} \rightarrow 2Al^{3+}{}_{(aq)} + 3Cu_{(s)}$

Identify the anode, cathode, direction of electron fl ow, migration of ions, and write the half-cell reactions.

8 Draw a diagram of the electrochemical cell represented by:

Mg | Mg^{2+} || Hg^{2+} | Hg, Pt

Use your knowledge of the reactivity of metals to determine which metal will reduce (displace) which metal ion. Hence write equations for the half-reactions and the overall reaction that occur in this cell. Indicate which electrode is positive and show the direction of electron flow and migration of ions, and identify the anode and cathode.

9 Explain the following:

a Dilute hydrochloric acid can be stored in copper containers, but not in aluminium ones.

10 Use the table of standard reduction potentials to answer the following:

a Rank the following in order of decreasing tendency to be reduced:

Cu^{2+}, Sn^{2+}, Ba^{2+}, Ag^{+}, Na^{+}

b Rank the following in decreasing tendency to be oxidised:

Cu, Mg, Au, Pb

11 What is the cell reaction and E° for a galvanic cell composed of the half-cells

Ni^{2+} | Ni || Fe^{3+}, Fe^{2+} | Pt

(You will need to refer to the table of standard reduction potentials.)

12 Write the anode and cathode half-cell equations and determine the EMF for the cell formed by linking the half-cells:

Cl^{-}, Cl_2 | Pt || I^{-}, I_2 | Pt

13 Write the half-reactions that make up the following overall reaction, then use a table of electrode potentials to decide whether or not the reaction occurs as written.

$MnO_4^{-} + 5Fe^{2+} + 8H^{+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

Source: Debra Smith, Conquering Chemistry HSC Course Blackline Masters, Worksheet 12, McGraw-Hill Australia, 2003.

WORKSHEET 1.9

ELECTROLYSIS

1. How does an electrolytic cell differ from a galvanic cell? (Consider energy required, spontaneity of chem. Reactions, processes occurring at electrodes and in solution, components of cells.)

2. What factors affect the reactions occurring within an electrolytic cell?

3. What factors affect the RATE of electrolysis within an electrolytic cell?

4. Show with a fully labeled diagram how electrolysis can be used to restore a corroded silver artifact (eg. Spoon). Describe the chemical processes occurring within your cell.

WORKSHEET 1.10

ELECTROLYSIS

1 Check your understanding by completing the following:

The process in which an electric current is used to bring about a chemical reaction is ___________________. A cell in which electrolysis occurs is called an ___________________ cell. In this type of cell a flow of electrons cause reduction at the cathode, which is the ___________________ electrode and __________________ at the anode, which is the __________________ electrode. An electrolyte solution is necessary to allow migration of the _______________ anions. Factors that determine the products of an electrolysis reaction are the

a chemical nature of the __________________ in the solution

b _______________ of the ions present

c _______________ of the electrodes.

The electrolytic process of depositing a thin film of metal on the surface of another object is called ____________________. The object to be plated is made the __________________ . The electrolyte is a solution containing the _________________ ion.

For Questions 2–8 circle the letter corresponding to the most correct answer.

2 Which of the following best identifies the anode in an electrolytic cell?

A The electrode at which anions are discharged

B The electrode at which no gas can be evolved

C The electrode at which hydroxide ions are produced

D The electrode at which oxidation occurs

3 Consider electrolysis of molten sodium chloride, using inert electrodes.

Which of the following equations represents the reaction at the positive electrode?

A $Na \rightarrow Na^+ + e^-$

B $Na^+ + e^- \rightarrow Na$

C $2Cl^- \rightarrow Cl_2 + 2e^-$

D $Cl_2 + 2e^- \rightarrow 2Cl^-$

4 When a dilute solution of hydrochloric acid undergoes electrolysis using inert electrodes:

A Oxygen is produced at the anode and chloride ions are oxidised

B Oxygen is produced at the cathode and chloride ions are oxidised

C Oxygen is produced at the anode and hydrogen ions are oxidised

D Hydrogen is produced at the anode and chloride ions are oxidised

5 The copper plating of an object involves electrolysis of a copper sulfate solution using a copper electrode. During this process:

A Copper metal is deposited on the positive electrode

B Hydrogen gas is given off at the negative electrode

C Copper ions migrate towards the anode

D The mass of the anode decreases

6 For an electrolytic cell the cathode is:

A Negative and the site of oxidation

B Positive and the site of oxidation

C Negative and the site of reduction

D Positive and the site of reduction

7 Which of the following does *not* affect the rate of an electrolytic reaction:

A Inert electrodes

B Voltage applied

C Concentration of ions in the electrolyte

D Distance between electrodes

8 If 500 electrons per second are being released at one electrode of an electrolytic cell, the number of electrons per second being used up at the other electrode is:

A At least 500

B Exactly 500

C Greater than 500

D Dependent on the chemicals used

9 For the electrolysis of molten magnesium bromide using graphite electrodes, predict the anode and cathode reactions, giving half-equations for them.

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10 For the electrolysis of a neutral nickel(II) chloride solution using inert platinum electrodes, predict the electrode reactions. Give the electrode half-reactions and the overall cell reaction.

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Metal Extraction Chemistry

11 For the electrolysis of a 1.00 mol/L aqueous solution of potassium sulfate using inert electrodes, predict the products at the anode and cathode, write the overall equation and determine the minimum cell voltage for the electrolysis to occur.

12 Predict the anode and cathode reactions for the electrolysis of copper bromide solution with copper electrodes.

13 Write the half-reactions for the electrolysis of molten potassium hydroxide. In the cell a silvery metal formed at the cathode and bubbles of gas were observed at the anode.

14 A neutral solution of $Pb(NO_3)_2$ is electrolysed using copper electrodes. Referring to the table of electrode potentials in your text book write the half-cell and overall cell reactions.

15 Complete the following table comparing galvanic and electrolytic cells:

Features	Galvanic cell	Electrolytic cell
Type of redox reaction		Non-spontaneous
Energy	Produced	
Anode reaction		
Cathode reaction		
Anode polarity		
Cathode polarity		
Electrolytic solution	Necessary	
Cation movement	Towards cathode	
Anion movement	Towards anode	
Structure		Usually single cell

16 a Give one example of electroplating.

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b Explain why electroplating is used in preference to other methods.

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c Draw a diagram to show how a steel bathroom tap could be plated with chrome. Include the anode, cathode, electrolyte and appropriate reactions.

Source: Debra Smith, *Conquering Chemistry HSC Course Blackline Masters*, Worksheet 4, McGraw-Hill Australia, 003.

EXPERIMENT 1.3

ANODISING ALUMINIUM

INTRODUCTION

In recent decades aluminium has replaced steel for many uses, such as door and window frames in buildings, guttering on houses and in many factory and office fittings. A main reason for this is that aluminium does not corrode. Aluminium is a passivating metal; this means it forms a protective, impervious oxide coating on the surface. This oxide coating protects the aluminium from further reaction.

Unfortunately this oxide coating is a dull, whitish colour, which is visually less appealing than the shiny aluminium metal. To make the oxidised aluminium more appealing in appearance it is often dyed to produce a brightly coloured product. The aluminium oxide layer will soak up dyes for permanent colouring, but to get a good, deep colour the oxide layer needs to be quite thick.

The process of producing a thick oxide layer is called anodising. For the anodising process to be effective, the piece of aluminium must be extremely clean. Once the cleaning process has been completed, do not touch the aluminium.

AIM To anodise and dye a piece of aluminium.

EQUIPMENT

- 2 × 250 mL beakers
- glass rod
- watch glass
- DC power source
- 2 electrical leads with alligator clips
- 2 × 10 cm lengths of copper wire
- lead sheet 2 cm × 5 cm with a hole in the top of the piece
- aluminium sheet 2 cm × 5 cm with a hole in the top of the piece
- fine steel wool
- wash bottle of distilled water
- 250 mL 1.5 mol/L sulfuric acid
- Dylon multipurpose clothes dyes (Kingfi sher Blue and Emerald Green)
- Quink permanent ink (blue, red, black)
- For the class using the fume cabinet:
- 200 mL 2 mol/L sodium hydroxide in a water bath at 50ºC
- 200 mL 3 mol/L nitric acid

SAFETY

Wear safety glasses and protective clothing. Sodium hydroxide, nitric acid and sulfuric acid are corrosive, so avoid contact with skin. If contact occurs, wash well with soap and water. Use gloves when cleaning the piece of lead and take care not to inhale the dust. The firrst part of the experiment (steps 3–5) must be performed in a fume cabinet. A light spray of sulfuric acid is produced during the anodising process so cover the beaker with a piece of paper towel.

PROCEDURE

1 Pass one of the pieces of copper wire through the hole in the piece of aluminium and twist it to secure. (Do not touch the piece of aluminium from now on; use the copper wire handle.)

2 Three-quarters fi ll a large beaker with distilled water.

3 In the fume cabinet dip the piece of aluminium in the 2 mol/L sodium hydroxide at 50ºC for about 10 seconds.

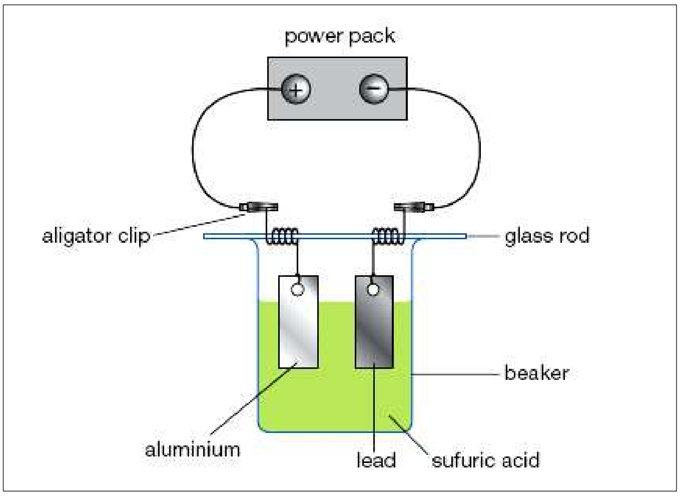
4 Rinse it with distilled water, then dip it in the 3 mol/L nitric acid for about 3 seconds.

5 Rinse again with distilled water, and place the piece of cleaned aluminium in a beaker of distilled water. Take care not to touch the cleaned aluminium.

6 Thread the other piece of copper wire through the hole in the piece of lead and twist it to secure it.

7 Use the steel wool to clean the piece of lead. Do not touch the cleaned lead.

8 Use the copper wire to secure the two pieces of metal to a glass rod, leaving enough wire to attach the alligator clips and lay the glass rod on top of the beaker. Ensure the two pieces of metal are not touching.



9 Carefully pour enough 1.5 mol/L sulphuric acid into the beaker to almost cover both pieces of metal. Leave about 1 cm above the surface.

10 Use the electrical leads with alligator clips to attach the piece of lead to the negative terminal and the piece of aluminium to the positive terminal.

11 Cover the beaker with a piece of paper towel, turn the power supply to 12 V, turn on the power and leave for 30 minutes. Record your observations.

12 After 30 minutes carefully remove the piece of aluminium using the copper wire (do not touch it), wash it with distilled water and store it in a beaker of distilled water.

13 Place the aluminium in a beaker of dye that has been immersed in a water bath at 50ºC. Leave for 10 minutes, gently moving the piece every minute.

14 Remove the strip from the dye and carefully wash it to remove excess dye.

15 Placed the aluminium in a beaker of boiled water for 10 minutes to seal the dye.

QUESTIONS

1 Why must the aluminium be thoroughly clean before it is anodised?

2 Why is it important that the aluminium is not touched after it has been cleaned?

3 a Identify the anode and cathode of the anodising cell.

b Write the half-cell reactions.

4 Suggest another metal that could have been used in place of lead.

EXPERIMENT 1.4

CLEANING TARNISHED SILVER

INTRODUCTION

Silver will react very slowly and tarnish over time. The tarnish that collects on objects made of silver is silver sulphide, a black solid. In the marine environment the formation of silver sulphide is often due to the action of sulfate reducing *Desulfovibrio* bacteria. It is also formed when silver forks and spoons are used with eggs or green vegetables, especially brussel sprouts.

One way to remove this black layer would be to use an abrasive to clean it, but this would take off some of the silver, as well as possibly damaging the engraving and embossing on the surface of the silver item. A preferred method is to clean the silver electrochemically so the Ag^+ ions in the silver sulphide are reduced to Ag.

In this investigation you will explore two ways of cleaning silver, both of which involve redox reactions.

AIM

To investigate two different techniques for cleaning tarnished silver and compare their effectiveness.

EQUIPMENT

- DC power source (0–12 V)
- 2 electrical leads with alligator clips
- 1 inert electrode, for example carbon (graphite), platinum or stainless steel
- 250 mL beaker
- 500 mL beaker or glass or plastic dish
- 200 mL 3 mol/L sodium hydroxide (NaOH) solution
- aluminium foil
- sodium hydrogen carbonate (NaHCO3)—approximately 2 teaspoons
- hot water
- items of tarnished silver, for example a spoon, a fork

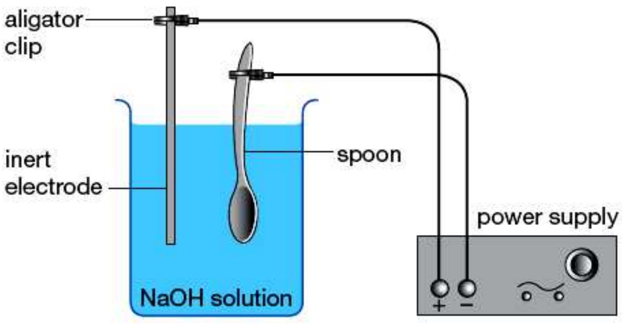
PROCEDURE

Part A Cleaning using electrolysis

1 Set up the apparatus as shown in the diagram.

2 Turn the current to 12 V and record your observations.

3 Allow the current to flow until the object is clean.



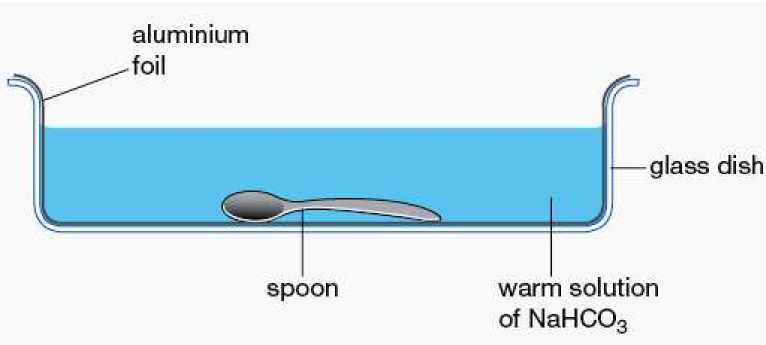
Part B Cleaning chemically

1 Line the beaker or dish with aluminium foil.

2 Place the object to be cleaned in the beaker or dish, ensuring it has maximum contact with the aluminium foil.

3 Add a heaped teaspoon of sodium hydrogen carbonate ($NaHCO_3$) and pour in enough hot water to ensure the object is covered.

4 Record your observations.



QUESTIONS

1 For part A:

a Identify the anode and cathode.

b Write the anode and cathode half reactions.

c Write the overall reaction.

2 For part B:

a Write the oxidation and reduction half reactions.

b Give two reasons why aluminium foil is used.

c Suggest what the black substance formed on the aluminium could be.

3 Why is a power supply needed for part A but not for part B?

a Why is an inert electrode used in part A?

b What effect (if any) would there be if the inert electrode were to be replaced with a silver electrode?

4 Compare the two objects that have been cleaned and comment on the effectiveness of the two methods used.