

CHEMISTRY SS2

Third term, 2022

William Ramsay.

"The country which is in advance of the rest of the world in chemistry will also be foremost in wealth and in general prosperity."

SS2 3rd Term Scheme of Work

WEEK	TOPIC	CONTENTS
1.	SULPHUR	<ul style="list-style-type: none"> - Electronic configuration, oxidation states, extraction, allotropes, properties, uses
2	SULPHUR	<ul style="list-style-type: none"> - Hydrogen sulphide; preparation, physical and chemical properties, test - Sulphur dioxide; preparation, properties
3	SULPHUR	<ul style="list-style-type: none"> - Sulphur trioxide; preparation, properties - Sulphuric acid; contact process, physical and chemical properties
4.	REDOX REACTIONS	<ul style="list-style-type: none"> - Definitions; oxidizing and reducing agents, tests - Oxidation number and connection to IUPAC name
5.	RADOX REACTIONS	<ul style="list-style-type: none"> - Redox equations and balancing
6.	IONIC THEORY	<ul style="list-style-type: none"> - Electrolytes and non-electrolytes. - Ionic theory. - Electrolysis and Mechanism of electrolytic cell, - Definition of terms, electrodes etc
7.	ELECTROLYSIS	<ul style="list-style-type: none"> - Factors affecting the preferential discharge of ions - Electrolysis of common electrolytes e.g. acidified water, brine etc
8.	ELECTROLYSIS	<ul style="list-style-type: none"> - applications of electrolysis, - Discuss electroplating, purification of copper and extraction of Aluminium
9	ELECTROLYSIS	<ul style="list-style-type: none"> - Faraday's laws of electrolysis; calculations
10	ELECTROLYSIS	<ul style="list-style-type: none"> - Electrochemical cells - Electrode potential; calculations - corrosion as an electrochemical process
11	RADIOACTIVITY	<ul style="list-style-type: none"> - Definition, types of radiation and characteristics, detection of radiation - Radioactive disintegration; alpha and beta decay, uses of radioisotopes - Artificial transmutation, nuclear fusion and fission and uses
12.	REVISION	<ul style="list-style-type: none"> - Revision
13.	EXAMINATION	<ul style="list-style-type: none"> - Examination

WEEK 1-3: SULPHUR AND ITS COMPOUNDS

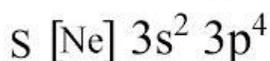
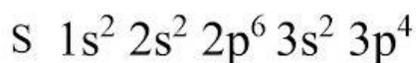
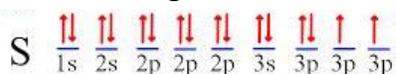
Sulphur is a chemical element with the symbol S and atomic number 16. It is abundant, multivalent, and nonmetallic. Under normal conditions, sulfur atoms form cyclic octatomic molecules with a chemical formula S₈.

atomic number: 16 [32.059, 32.076]
 symbol: S
 electron configuration: [Ne]3s²3p⁴
 name: sulfur*

atomic weight: [32.059, 32.076]
 acid-base properties of higher-valence oxides: Strongly acidic
 crystal structure: Orthorhombic
 physical state at 20 °C (68 °F): Solid

Other nonmetals	Solid
Orthorhombic	Strongly acidic

Electronic configuration



Oxidation numbers

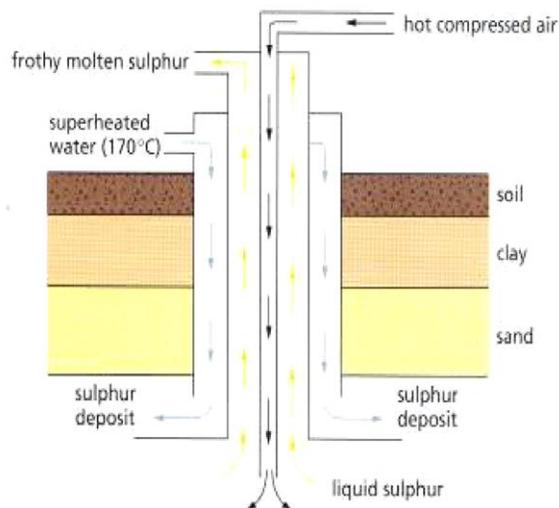
Chemical formulae	Oxidation state
S ²⁻ , H ₂ S	-2
S ₈	0
SCl ₂ , S ₂ O ₃ ²⁻	+2
SO ₂ , SO ₃ ²⁻ , H ₂ SO ₃	+4
SO ₃ , SO ₄ ²⁻ , H ₂ SO ₄	+6

Extraction

The Frasch process

Superheated water at 170°C and hot compressed air are forced underground through pipes, forcing water and molten sulphur to the surface. Sulphur is insoluble in water and so the two substances emerging from the pipes are easily separated. The sulphur is kept molten and sold

in this form. The sulphur obtained from this process is about 99.5% pure and can be used directly.



Allotropes of sulphur

1. **Crystalline** allotropes
 - i. Rhombic sulphur (α -sulphur)
 - ii. Monoclinic or prismatic sulphur. (β -sulphur)
2. **Amorphous** sulphur (δ -sulphur).

There is another form known as plastic sulphur

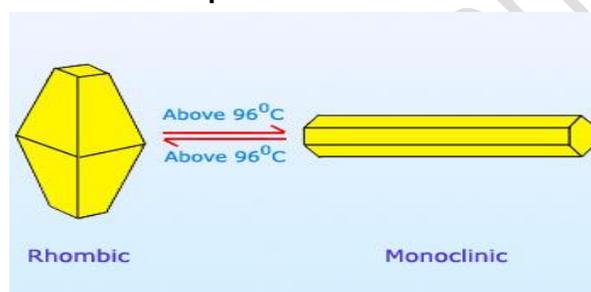
Rhombic sulphur is a crystalline allotropic form of sulphur. It is also designated as α -sulphur. This is the most stable variety of sulphur. All other varieties of sulphur eventually revert to rhombic form on standing.

Monoclinic sulfur is a crystalline allotrope of sulfur obtained when rhombic sulfur is heated to 94.5°C. This form is stable only above 96°C.

When left at room temperature it reverts back to rhombic form. It has S₈ ring molecules in crystalline structure.



Rhombic sulphur is stable below 96°C and monoclinic sulphur is stable above 96°C. This temperature is called the **transition temperature**.



Comparing the physical properties of rhombic and monoclinic sulphur

Properties	Rhombic	Monoclinic
Colour	Bright yellow	amber
Shape	octahedral	Needle-shaped
Density (g/cm ³)	2.08	1.98
Melting point	113°C	119°C
Stability	Stable below 96°C	Stable between 96°C and 119°C

Amorphous sulphur; is prepared as a pale yellow almost white deposit when hydrogen sulphide is bubbled through water for a long time and the saturated solution is exposed to air. It can also be deposited by the action of dilute hydrochloric acid on trioxothiosulphate (VI) solution

Plastic sulphur; this is prepared by heating and melting solid sulphur and the resultant liquid is poured into a beaker of cold water. This is an elastic, rubber-like substance and will not dissolve in carbon (IV) sulphide.



Physical Properties of Sulphur

1. It is a yellow, brittle solid at room temperature.
2. Is insoluble in water but soluble in carbon (IV) sulphide and methyl benzene (toluene).
3. It is a bad conductor of heat and electricity
4. It has a melting point of 119°C and a boiling point of 444°C

Chemical properties of sulphur

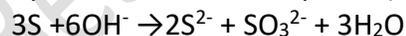
1. Sulphur will react with both metals and non-metals

$$\text{Mg(s)} + \text{S(s)} \rightarrow \text{MgS(s)}$$

$$\text{S(s)} + \text{O}_2\text{(g)} \rightarrow \text{SO}_2\text{(g)}$$
2. Sulphur is readily oxidized when warmed with concentrated sulphuric acid to form sulphur (IV) oxide. When warmed with conc. Nitric acid, using bromine as a catalyst sulphur is oxidized to sulphuric acid

$$2\text{H}_2\text{SO}_{4\text{(aq)}} + \text{S(s)} \rightarrow 2\text{H}_2\text{O(l)} + 3\text{SO}_2\text{(g)}$$

$$6\text{HNO}_{3\text{(aq)}} + \text{S(s)} \rightarrow \text{H}_2\text{SO}_{4\text{(aq)}} + 6\text{NO}_2\text{(g)} + 2\text{H}_2\text{O(l)}$$
3. Sulphur will react with hot concentrated alkaline solutions to form a mixture of sulphides and trioxosulphates (IV) which in the presence of excess sulphur react to form a polysulphide and a trioxosulphate (VI) respectively.



Assignment 1

1. State 4 uses of sulphur
2. What is vulcanisation?

Hydrogen Sulphide, H₂S

This is prepared both in the laboratory and commercially by the action of dilute acid on a metallic sulphide like iron (II) sulphide FeS



Physical properties of hydrogen sulphide

1. Colourless gas with a repulsive smell like that of a rotten egg
2. Very poisonous

- 1.18 times denser than air
- Moderately soluble in water to form a weak acidic solution

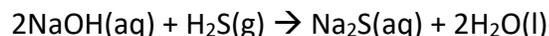
Chemical properties

- The major chemical property of this gas is that it's a strong reducing agent and would react with sulphuric acid and other oxidizing agents to deposit sulphur



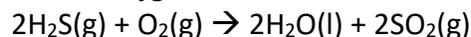
(Make reference to the new school chemistry textbook for reactions with other oxidizing agents)

- As an acid it reacts with alkali to form salt and water.

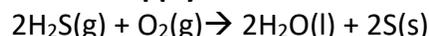


- With oxygen; it burns in air with a bright blue flame

Excess oxygen

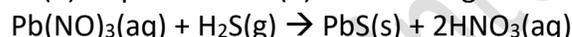


limited supply



Test

To test for the gas, moisten a piece of filter paper with lead (II) trioxonitrate (V) solution and drop it in a jar of the unknown gas if the gas is H₂S, the paper turns black due to the deposition to the formation of lead (II) sulphide. Lead (II) ethanoate gives the same reaction.



Sulphur (IV) Oxide, SO₂.

It is a toxic gas responsible for the smell of burnt matches

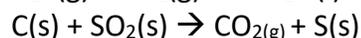
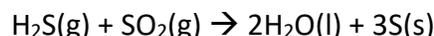
Laboratory Preparation

This is an acid anhydride and is prepared in the laboratory by heating sodium or potassium tetraoxosulphate (VI) with tetraoxosulphate (VI) or hydrochloric acid. The trioxosulphate (IV) acid formed breaks down to yield water and sulphur (IV) oxide.

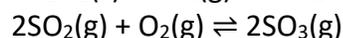
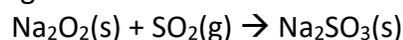


Chemical properties

- As an acid anhydride it reacts with alkalis to form salt and water only
- As a reducing agent; it decolourises KMnO₄, changes K₂Cr₂O₇ from orange to green and Brown Fe³⁺ to green Fe²⁺
- As an oxidizing agent; in the presence of stronger reducing agent it acts as an oxidizing agent



- It undergoes direct combination with certain metallic oxides and with oxygen in the following manner



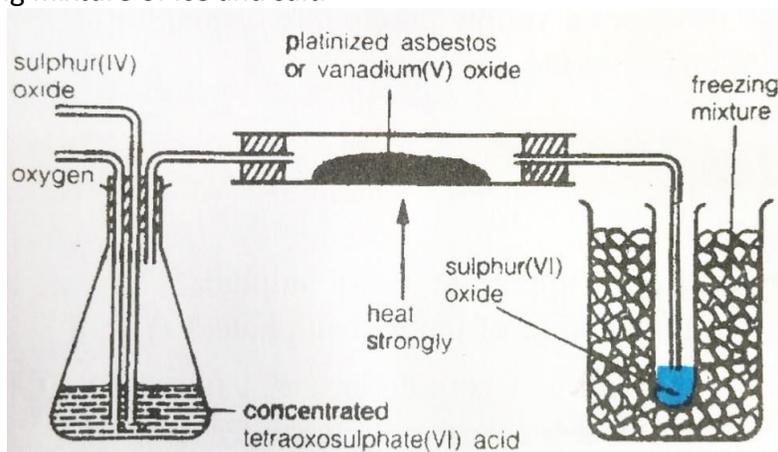
- It is a bleaching agent

Assignment 2

1. Draw and write a short note on the kipp's apparatus
2. Differentiate between the bleaching action of SO₂ and Cl₂

Sulphur (VI) oxide, SO₃

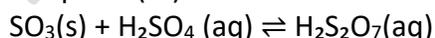
In the laboratory, a mixture of sulphur(IV) oxide and oxygen is first dried by passing it through concentrated tetraoxosulphate(VI) acid. Then the gaseous mixture is passed over a strongly heated catalyst. The sulphur(VI) oxide formed is seen as dense white fumes. It solidifies when cooled in a freezing mixture of ice and salt.



Sulphur (VI) oxide can also be obtained by dehydrating tetraoxosulphate(VI) acid with phosphorus(V) oxide, or by heating some tetraoxosulphates(VI) or hydrogentetraoxo sulphates(VI).

Properties

1. At room temperature, sulphur(VI) oxide exists as white needle-shaped crystals. It has a low boiling point (45°C) and hence readily vaporizes on gentle heating.
2. It is an acidic oxide. It reacts exothermically with water to form tetraoxosulphate (VI) acid, and neutralizes bases to form tetraoxosulphates (VI). Hence, it is also known as the tetraoxosulphate(VI) anhydride.
3. It also dissolves in concentrated tetraoxo sulphate (VI) acid to form oleum or fuming tetraoxosulphate(VI) acid.



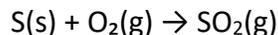
Tetraoxosulphate (VI) acid, H₂SO₄

This is one of the most important chemical compounds known. Almost every manufacturing process makes use of this acid directly or indirectly at some stage.

The Contact Process

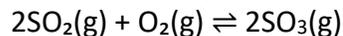
There are three stages in the Contact process.

Stage 1: The oxidation of sulphur to sulphur dioxide in the furnace, sulphur is burnt in a stream of dry purified air. Here, sulphur is oxidised by the oxygen in the air to form sulphur dioxide.



Stage 2: The oxidation of sulphur dioxide to sulphur trioxide

This stage occurs in the converter. Under controlled conditions, sulphur dioxide is reacted with oxygen to form sulphur trioxide. The process is exothermic

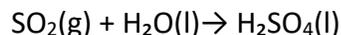


The condition for this stage are

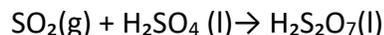
- i. A catalyst like platinised asbestos or **vanadium (v) oxide**, V_2O_5
- ii. A slight pressure and
- iii. A temperature of 400-450 °C

Stage 3: Conversion of sulphur trioxide to sulphuric (VI) acid

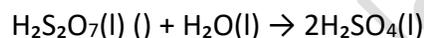
Sulphur trioxide could be dissolved in water to give sulphuric (VI) acid.



However, this reaction is too violent and dangerous. It gives out large amounts of heat and produces a choking mist of sulphuric (VI) acid. To avoid this, sulphur trioxide is dissolved concentrated sulphuric (VI) acid to form a fuming liquid called oleum.



Oleum is then carefully diluted with the appropriate quantity of water to make concentrated sulphuric(VI) acid.

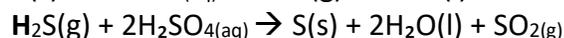


Physical Properties

1. Concentrated tetraoxosulphate(VI) acid, often called the oil of vitriol is a colourless, viscous liquid with a density of 1.84 g cm^3 ,
2. It is corrosive and causes severe burns if it comes into contact with the skin.
3. Concentrated tetraoxosulphate(VI) acid has a great affinity for water, evolving a large amount of heat as it dissolves. This heat evolved is a result of the hydration of the tetraoxosulphate(VI) ions. The concentrated acid is also hygroscopic, absorbing water vapour from the surroundings and becoming more dilute in the process.

Chemical Properties

1. As an acid, it reacts with metals, bases and trioxocarbonates
2. as an oxidising agent ; the concentrated acid oxidizes metals, non-metals and hydrogen sulphide to produce sulphur oxide and other corresponding compounds



3. it is also a dehydrating agent

Assignment 3

1. Differentiate between the drying and dehydrating action of sulphuric acid using appropriate examples
2. State 10 uses of sulphuric acid

Week 4-5: Redox Reactions

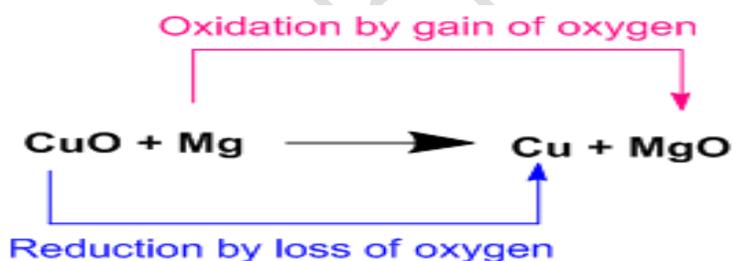
What are Redox reactions?

Redox is a type of chemical reaction in which the oxidation states of atoms are changed. Redox reactions are characterized by the actual or formal transfer of electrons between chemical species, most often with one species undergoing oxidation while another species undergoes reduction

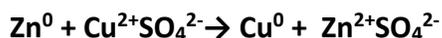
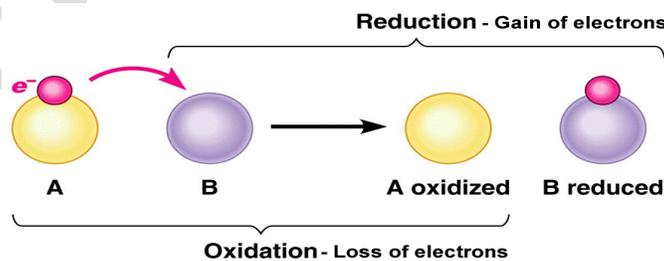
How then do we define oxidation and reduction?

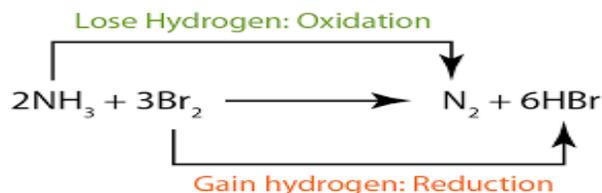
There are three major ways we could define them; by loss or gain of

1. oxygen
 2. hydrogen
 3. electrons
- For oxygen, oxidation is gain and reduction is loss (of oxygen)

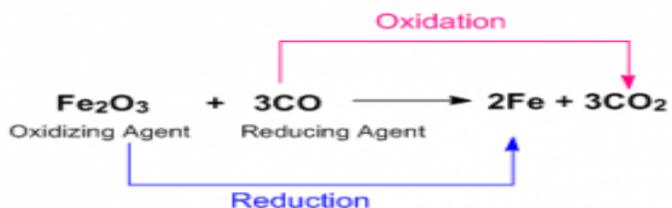


- For hydrogen and electron (memory aid; OIL RIG): oxidation is loss and reduction is gain





What are oxidizing and reducing agents?



**REDUCING AGENT
VERSUS
OXIDIZING AGENT**

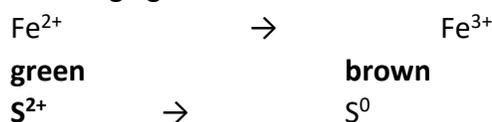
A reducing agent is a substance that can be oxidized by losing some of its electrons	An oxidizing agent is a substance that can be reduced by obtaining electrons
Oxidation state increases	Oxidation state decreases
Acts as the electron donor	Acts as the electron receiver
Oxidized during the reaction	Reduced during the reaction
Causes the reduction of another reactant	Causes the oxidation of another reactant
	Visit www.psdiaa.com

Test for oxidizing and reducing agents

- **Some oxidizing agents include**
 - Oxygen, O₂
 - manganese (IV) oxide, MnO₂
 - Acidified KMnO₄
 - acidified K₂Cr₂O₇ etc
- **Some reducing agents include**
 - carbon, C
 - Hydrogen, H₂
 - Carbon(II) oxide, CO
 - Iron (II) salt, Fe²⁺
 - Hydrogen sulphide, H₂S e.t.c

- **Test for oxidizing agents, OA**

the action of oxidizing agents on iron (II) chloride and hydrogen sulphide is frequently employed in the qualitative analysis as test for the identification of oxidizing agents

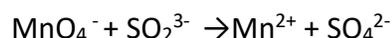


- **Test for Reducing agents, RA**

reducing agents change a solution of acidified potassium tetraoxomanganate (VII), KMnO_4 , from **purple** to colourless and acidified potassium dichromate, $\text{K}_2\text{Cr}_2\text{O}_7$ from **orange** to **green**

Exercise (more from other sources)

Identify the oxidizing agent and the reducing agent in the following Redox reaction:



How does IUPAC naming relate to oxidation number, ON?

As an example; find the oxidation number of the chromium atom in $\text{K}_2\text{Cr}_2\text{O}_7$

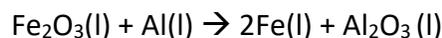
Where ON of K = +1, Cr = x and O = -2

$$\begin{aligned} 2(\text{K}) + 2(\text{Cr}) + 7(\text{O}) &= 0 \\ 2(+1) + 2(x) + [7 \times (-2)] &= 0 \\ 2 + 2x - 14 &= 0 \\ 2x &= +12 \\ x &= +6 \end{aligned}$$

The oxidation number of chromium atom is +6. Therefore the name of the compound is potassium heptaoxodichromate (VI)

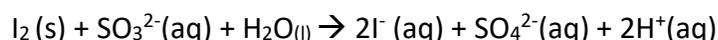
Assignment 4

1. A redox reaction occurs when molten aluminum reacts with iron (III) oxide. Indicate each of the following for the preceding redox reaction:



(a) substance oxidized (b) substance reduced (c) oxidizing agent (d) reducing agent

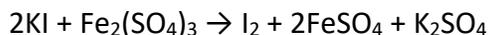
2. The amount of iodine in a solution can be determined by a redox method using a sulfite solution:



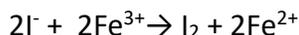
Indicate each of the following for the preceding reaction: (a) substance oxidized (b) substance reduced (c) oxidizing agent (d) reducing agent

So how do I identify a redox reaction?

- The first step in balancing any redox reaction is determining whether or not it is even an oxidation-reduction reaction, which requires that species exhibits changing oxidation states during the reaction.
- To maintain charge neutrality in the sample, the redox reaction will entail both a reduction component and an oxidation components



Ionically



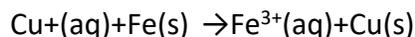
How do I go about balancing a redox equation?

The method used to balance redox reactions is called the Half Equation Method. In this method, the equation is separated into two half-equations; one for oxidation and one for reduction. Each equation is balanced by adjusting coefficients and adding H_2O , H^+ , and e^- in this order:

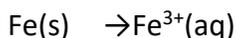
1. Write the half equations
2. Balance elements in the equation other than O and H.
3. Balance the oxygen atoms by adding the appropriate number of water (H_2O) molecules to the opposite side of the equation.
4. Balance the hydrogen atoms (including those added in step 2 to balance the oxygen atom) by adding H^+ ions to the opposite side of the equation.
5. Add up the charges on each side. Make them equal by adding enough electrons (e^-) to the more positive side. (Rule of thumb: e^- and H^+ are almost always on the same side.)
6. The e^- on each side must be made equal; if they are not equal, they must be multiplied by appropriate integers (the lowest common multiple) to be made the same.
7. The half-equations are added together, canceling out the electrons to form one balanced equation. Common terms should also be canceled out.

Now let us balance an equation

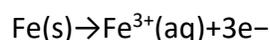
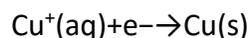
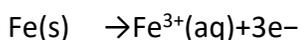
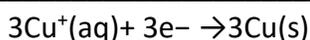
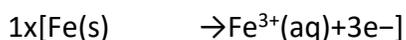
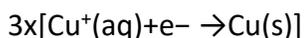
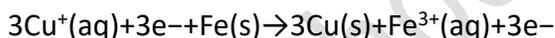
Example 1



Solution

Step 1:**oxidation half****reduction half****Step 2:** nil**Step 3:** nil

Step 4: Balance the electrons in the equations. In this case, the electrons are simply balanced by multiplying the entire half-reaction by 3 and leaving the other half reaction as it is. This gives

**Step 5:****Step 6:** Adding the equations give:

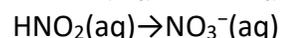
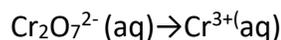
The electrons cancel out and the balanced equation is left.

**Exercise**

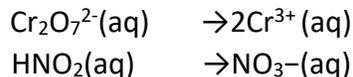
balance the redox equation

**Example 2**

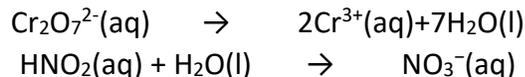
- Step 1: Separate the half-reactions. The table provided does not have acidic or basic half-reactions, so just write out what is known.



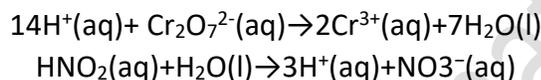
- Step 2: Balance elements other than O and H. In this example, only chromium needs to be balanced. This gives:



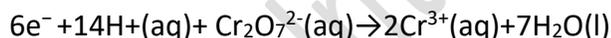
- Step 3: Add H₂O to balance oxygen. The chromium reaction needs to be balanced by adding 7 H₂O molecules. The other reaction also needs to be balanced by adding one water molecule. This yields:



- Step 4: Balance hydrogen by adding protons (H⁺). 14 protons need to be added to the left side of the chromium reaction to balance the 14 (2 per water molecule by 7 water molecules) hydrogen. 3 protons need to be added to the right side of the other reaction.



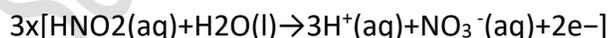
- Step 5: Balance the charge of each equation with electrons. The chromium reaction has (+14) + (-2) = +12 on the left side and (2 x +3) = +6 on the right side. To balance, add 6 electrons (each with a charge of -1) to the left side:



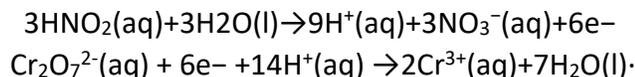
- For the other reaction, there is no charge on the left and a (+3) + (-1) = +2 charge on the right. So add 2 electrons to the right side:



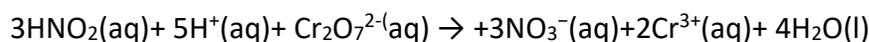
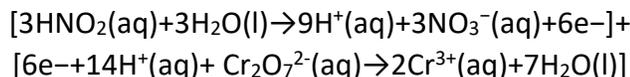
- Step 6:** Scale the reactions so that the electrons are equal. The chromium reaction has 6e⁻ and the other reaction has 2e⁻, so it should be multiplied by 3. This gives:



it becomes



- Step 7:** Add the reactions and cancel out common terms.



The electrons cancel out as well as 3 water molecules and 9 protons/hydrogen ion.

Exercise/assignment 5

Task: Balance the following redox reactions:

1. $\text{SO}_2 (\text{g}) + \text{HNO}_2 (\text{aq}) \rightarrow \text{H}_2\text{SO}_4 (\text{aq}) + \text{NO} (\text{g})$
2. $\text{Al} (\text{s}) + \text{H}_2\text{SO}_4 (\text{aq}) \rightarrow \text{Al}_2(\text{SO}_4)_3 (\text{aq}) + \text{H}_2 (\text{g})$
3. $\text{Au}^{3+} (\text{aq}) + \text{I}^- (\text{aq}) \rightarrow \text{Au} (\text{s}) + \text{I}_2 (\text{s})$
4. $\text{S}^{2-} (\text{aq}) + \text{I}_2 (\text{s}) \rightarrow \text{SO}_4^{2-} (\text{aq}) + \text{I}^- (\text{aq})$
5. $\text{Br}_2 (\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{Br}^- (\text{aq}) + \text{BrO}_3^- (\text{aq})$
6. $\text{H}_2\text{O}_2 (\text{aq}) + \text{ClO}_4^- (\text{aq}) \rightarrow \text{O}_2 (\text{g}) + \text{ClO}_2^- (\text{aq})$
7. $\text{Mn} (\text{s}) + \text{HNO}_3 (\text{aq}) \rightarrow \text{Mn}^{2+} (\text{aq}) + \text{NO}_2 (\text{g})$
8. $\text{I}_2 (\text{s}) + \text{OCl}^- (\text{aq}) \rightarrow \text{IO}_3^- (\text{aq}) + \text{Cl}^- (\text{aq})$
9. $\text{Cr}_2\text{O}_7^{2-} (\text{aq}) + \text{HNO}_2 (\text{aq}) \rightarrow \text{Cr}^{3+} (\text{aq}) + \text{NO}_3^- (\text{aq})$
10. $\text{CrO}_4^{2-} (\text{aq}) + \text{S}^{2-} (\text{aq}) \rightarrow \text{Cr}(\text{OH})_3 (\text{s}) + \text{S} (\text{s})$

Week 6: The ionic theory

It is assumed that solid electrolyte consist of two types of charged particles, one carrying a positive charge and another carrying a negative charge. They are held together by the electrostatic force of attraction. When such solid are dissolves in a solvent, these forces are weakened and the electrolyte undergoes dissociation into ions. This process is called ion solvation.

The ionic theory is used to explain the process of electrolysis in an electrolyte, the ions are free to move in all directions. When an electric current is passed through the electrolyte, the negative ions are attracted to the positive and they move in its direction while the positive ions are attracted to the negative electrodes and move towards it.

Electrolyte and non-electrolytes

Electrolytes: these are molten ionic compounds or aqueous solutions containing ions which can conduct electricity and are decomposed in the process.

Strong electrolytes: Substances which dissociate almost completely in their aqueous solutions even at moderate dilutions are called strong electrolytes. Their dissociation reaction is irreversible.

Examples:

- Strong acids like HCl, HNO₃, H₂SO₄ etc.
- Salts like NaCl, KCl,
- Substances like H₂S etc.

Weak electrolytes: Substances which dissociate to a little (limited) extent in their aqueous solutions are called weak electrolytes.

Examples:

- All weak acids like CH₃COOH, HCOOH,
- All weak bases like NH₂OH,
- Salts like CH₂COONH₄, CH₃COOAg etc.

Non-electrolyte: this is a substance which in its aqueous solution or in the fused state does not conduct electricity (due to no formation of ions). Examples: sugar, urea, ethanol, starch, acetone, etc.

Assignment 6

1. State 4 differences between a covalent and an electrovalent compound
2. State how concentration and degree of ionization affect the conductance of an acid, base and salt.

Week 7-10: Electrolysis

Electrolysis is defined as the chemical decomposition produced by passing an electric current through a liquid or solution containing ions.

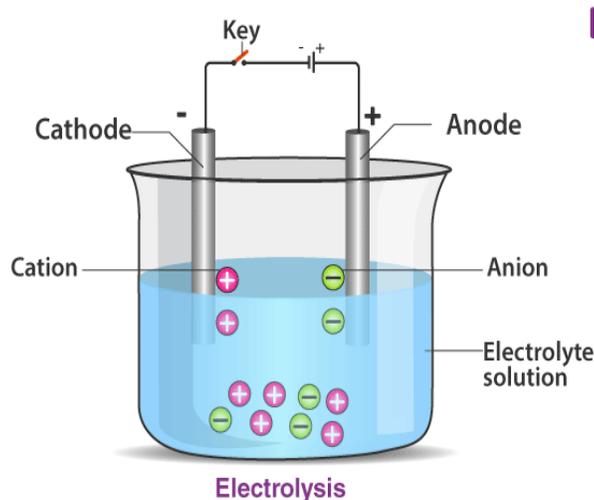
Mechanism of Electrolysis

Before the passage of current, the ions move about randomly in the electrolyte. When electrolysis begins cations are attracted to the negatively charged cathode where they accept electrons and the anions are attracted to the positively charged anode where they give up their electrons. Both ions become electrically neutral and are discharged.

Electrode; these are conductors in the form of wires, rods or plates through which an electric current enters or leaves the electrolyte.

Anode: this is the electrode where oxidation occurs

Cathode: this is the electrode where reduction occurs



Factors that affect preferential discharge of ions during electrolysis

1. position of ions in the electrochemical series
2. concentration of ions in the electrolyte
3. the nature of electrode

Position of ions in the electrochemical series; ions lower in the electrochemical series are preferentially discharge to the ones higher in the series

The Electrochemical Series

K ⁺	↓ Ease of discharge increases	F ⁻
Na ⁺		SO ₄ ²⁻
Ca ²⁺		NO ₃ ⁻
Mg ²⁺		Cl ⁻
Al ³⁺		Br ⁻
Zn ²⁺		I ⁻
Fe ²⁺		OH ⁻
Sn ²⁺		
Pb ²⁺		
H ⁺		
Cu ²⁺		
Ag ⁺		

(do you remember the memory aid for the electrochemical series?)

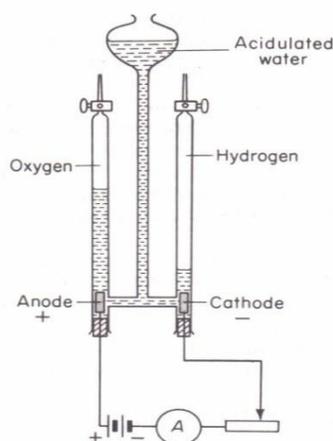
Concentration of ions; If other conditions are equal, increasing the concentration of a given ion tends to promote its discharge from solution. The influence of concentration, however, is effective only when the two competing ions are closely positioned in the electrochemical series.

The effect of concentration becomes less important as the positions of the competing ions become further apart in the series.

For example, in concentrated sodium chloride solution (i.e. brine), the two anions present are the chlorine ion and the hydroxyl ion. Although the hydroxyl ion is more easily oxidized than the chlorine ion, it is the chlorine ion which will be discharged because its concentration is much greater than that of the hydroxyl ion.

Electrolysis of common electrolytes

Electrolysis of acidified water; this is usually carried out in the laboratory using the Hofmann's Voltammeter



parameters	At the cathode H^+	At the anode OH^-, SO_4^{2-}
Ions discharged	$2H^+ + 2e \rightarrow H_2$	$4OH^- \rightarrow 2H_2O + O_2 + 4e$
Observation	Effervescence; a colourless gas at both the anode and the cathode	
Changes in electrolyte	The electrolyte becomes concentrated	
Over all equation	$2 \times (2H^+ + 2e \rightarrow H_2)$ $4OH^- \rightarrow 2H_2O + O_2 + 4e$ <hr/> $4H^+ + 4OH^- + 4e \rightarrow 2H_2O + O_2 + 4e$ $\text{Or } 4H_2O \rightarrow 2H_2O + O_2$	

Electrolysis of Dilute Sodium Chloride, $NaCl(aq)$

parameters	At the cathode H^+, Na^+	At the anode OH^-, Cl^-
Ions discharged	$2H^+ + 2e \rightarrow H_2$	$4OH^- \rightarrow 2H_2O + O_2 + 4e$
Observation	Effervescence; a colourless gas at both the anode and the cathode	
Changes in electrolyte	The electrolyte becomes concentrated	
Over all equation	$2 \times (2H^+ + 2e \rightarrow H_2)$ $4OH^- \rightarrow 2H_2O + O_2 + 4e$ <hr/> $4H^+ + 4OH^- + 4e \rightarrow 2H_2O + O_2 + 4e$ $\text{Or } 4H_2O \rightarrow 2H_2O + O_2$	

Electrolysis of concentrated sodium chloride, brine

parameters	At the cathode H^+, Na^+	At the anode OH^-, Cl^-
Ions discharged	$2H^+ + 2e \rightarrow H_2$	$2Cl^- \rightarrow 2Cl_2 + 2e$
Observation	Effervescence at both the anode and the cathode; a colourless gas at cathode but a greenish-yellow gas at anode	
Changes in electrolyte	The electrolyte becomes alkaline	
Over all equation	$2H^+ + 2e \rightarrow H_2$ $2Cl^- \rightarrow 2Cl_2 + 2e$ <hr/> $2H^+ + 2Cl^- \rightarrow H_2 + 2Cl_2$	

Electrolysis of molten sodium chloride, NaCl (l)

parameters	At the cathode Na^+	At the anode Cl^-
Ions discharged	$Na^+ + e \rightarrow Na$	$2Cl^- \rightarrow 2Cl_2 + 2e$

Observation	Effervescence at the anode greenish-yellow gas evolved. The cathode becomes larger as sodium is deposited
Changes in electrolyte	The electrolyte depletes
Over all equation	$2x(\text{Na}^+ + e \rightarrow \text{Na})$ $2\text{Cl}^- \rightarrow 2\text{Cl}_2 + 2e$ <hr/> $2\text{Na}^+ + 2\text{Cl}^- \rightarrow 2\text{Na} + 2\text{Cl}_2$ $\text{or } 2\text{NaCl} \rightarrow 2\text{Na} + 2\text{Cl}_2$

Assignment 7

As you have been taught, describe the processes in the electrolysis of

1. copper sulphate solution using inert electrodes
2. molten zinc bromide using inert electrodes

Application of electrolysis

1. Purification of metals
2. Electroplating of one metal by another
3. Extraction of elements e.g aluminium, Sodium, calcium etc
4. Preparation of certain important compounds such as sodium hydroxide

Effect of nature of electrode on preferential discharge of ions.

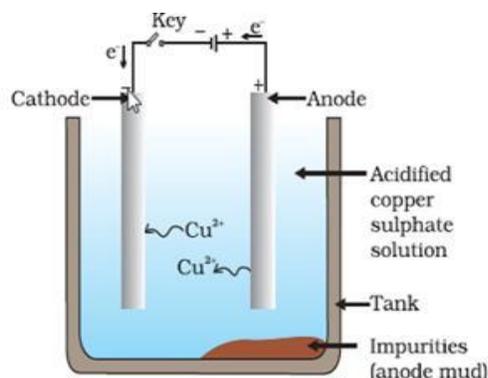
1. There two types of electrodes;
 - I. **Inert** electrodes e.g. graphite/carbon and platinum, although graphite is attacked by oxygen and platinum by chlorine
 - II. **Active** electrode e.g. copper, zinc, iron etc.

When an active electrode is made the anode and immersed into a solution of its salt during electrolysis, none of the anions present are discharged instead the active anode goes into solution and is discharged.

2. Some electrodes which have strong affinity for certain ions may influence ionic discharge. For example, in the electrolysis of aqueous sodium chloride using platinum electrodes, H^+ are preferentially discharged. If a mercury cathode is used, however, the mercury will tend to associate with the Na to form sodium **amalgam**, Na/Hg, so that the discharge of Na^+ requires less energy than that of H^+ , and so occurs in preference.

Purification of Copper by Electrolysis

In this process the metal to be purified serves as the anode while and its pure form serve as cathode. Both electrodes are inserted into a solution of its salt



parameters	At the cathode (pure/blister copper) H^+, Cu^{2+}	At the anode (impure copper) OH^-, SO_4^{2-}
Ions discharged	$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu(s)$	$Cu(s) \rightarrow Cu^{2+} + 2e^-$
Observation	Brown deposit of copper around the cathode and the anode reduces in size	
Changes in electrolyte	The blue electrolyte remains unchanged	
Over all equation	$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu(s)$ $Cu(s) \rightarrow Cu^{2+} + 2e^-$ <hr/> $Cu^{2+}_{(aq)} + Cu(s) \rightarrow Cu^{2+}(aq) + Cu(s)$	

Electroplating

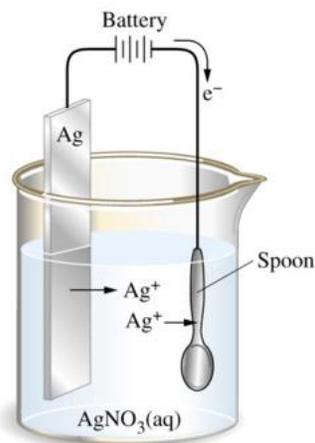
Electroplating is a method of coating the surface of one metal with another metal, usually copper, silver, chromium, nickel or gold, by means of electrolysis, for decoration or protection against corrosion.

The cathode is the metallic object which is to be coated or plated; the anode, a strip of the plating metal; and the electrolyte, a salt of the plating metal. Thus, in silver-plating a spoon,

- The cathode is the spoon.
- The anode is a silver rod, and

- The electrolyte is a soluble silver salt, usually a silver trioxonitrate (V) solution.

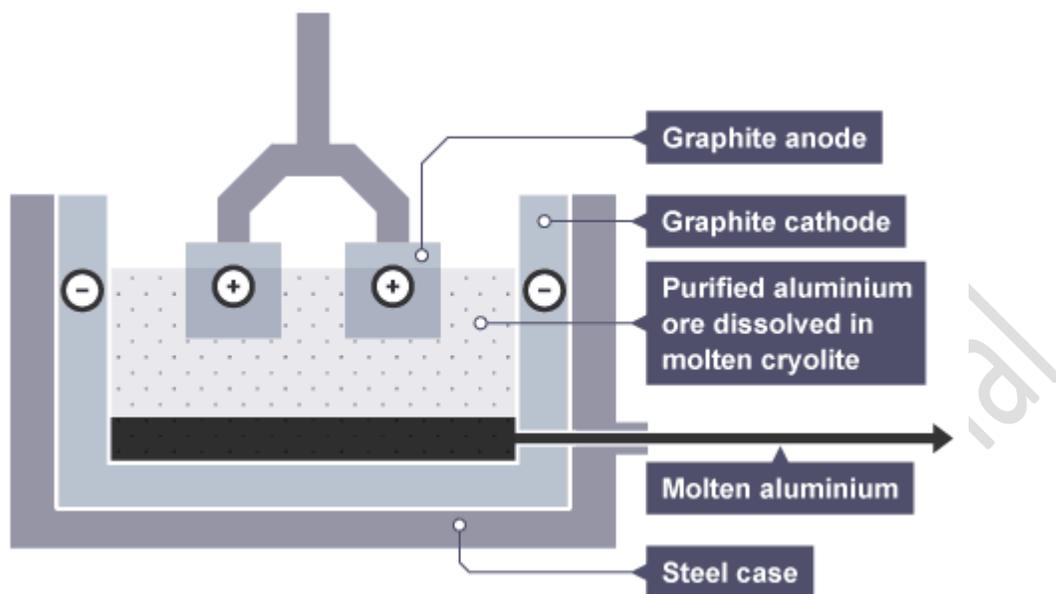
As the current is passed through the cell, plating metal dissolves at the anode, and the ions produced migrate to the cathode, where they are discharged and deposited as a layer on the object.



Extraction of aluminium

Aluminium ore is called bauxite. The **bauxite** is purified to produce aluminium oxide (**alumina**), a white powder from which aluminium can be extracted.

The extraction is done by electrolysis. The ions in the aluminium oxide must be free to move so that electricity can pass through it. Aluminium oxide has a very high melting point (over 2000°C) so it would be expensive to melt it. Aluminium oxide does not dissolve in water, but it does dissolve in molten cryolite. This is an aluminium compound with a lower melting point than aluminium oxide. The use of cryolite reduces some of the energy costs involved in extracting aluminium.



During electrolysis:

- positively charged aluminium ions gain electrons from the **cathode**, and form molten aluminium

$$\text{Al}^{3+}(\text{l}) + 3\text{e}^- \rightarrow \text{Al}(\text{l})$$
- oxide ions lose electrons at the **anode**, and form oxygen molecules

$$2\text{O}^{2-}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{e}^-$$
- the overall reaction

$$2\text{Al}_2\text{O}_3(\text{l}) \rightarrow 4\text{Al}(\text{l}) + 3\text{O}_2(\text{g})$$
- What happens to the anode after sometime?
 The oxygen reacts with the carbon in the electrodes, forming carbon dioxide which bubbles off. Carbon is therefore lost from the positive electrodes, so they must be replaced frequently. This adds to the cost of the process.

$$\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$$

Assignment 8

Electrolysis is used to electroplate the heat-reflecting shields of space rockets with gold. With the aid of a diagram and equations explain how electrolysis can be used to gold plate the heat reflecting shields.

Faraday's Laws of Electrolysis

Faraday's first law of electrolysis

The mass (m) of an element discharged during electrolysis is directly proportional to the quantity of electricity (Q) passing through it

$$Q=It$$

$$m = \frac{Q \times ram}{F \times Z}$$

Where

Q= quantity of electricity

I= current

t= time in seconds

m=mass in grams

F= faradays constant (96,500)

ram= relative atomic mass of element

Z= charge of element e.g. $Al^{3+} = 3$

Example

1. Calculate the quantity of electricity passed when 0.4 A flows for 1 hour 20 minutes through an electrolytic cell

$$Q=It$$

$$I= 0.4A$$

$$t= [(1 \times 60 \times 60) + (20 \times 60)]s \\ = 4800s$$

$$Q= 0.4A \times 4800s \\ = 19200C$$

2. What mass of copper would be formed when a current of 10.0 A is passed through a solution of $CuSO_4$ for 1 hour ($Cu=63.5$, $F= 96500C$)

Solution

$$m = \frac{Q \times ram}{F \times Z}$$

$$Q=It= 3600 \times 10.0 A \\ = 36000C$$

$$ram= 63.5$$

$$F= 96500$$

$$Z=2$$

Therefore

$$m = \frac{36000 \times 63.5}{96500 \times 2}$$

$$m = 11.8g$$

Faraday's second law of electrolysis

When the same quantity of electricity is passed through two different electrolytes, the relative number of moles of element discharged are inversely proportional to the charges on the ions of the element

$$\frac{m_1 \times Z_2}{m_2 \times Z_1} = \frac{m_2 \times Z_1}{m_1 \times Z_2}$$

Example

A voltameter containing silver trioxonitrate (V) solution was connected in series to another voltameter containing copper (II) tetraoxosulphate (VI) solution. When a current of 0.2 A was passed through the solutions, 0.780g of silver was deposited. Calculate the

- mass of copper that would be deposited in the copper voltameter
 - Quantity of electricity used and the time of current flow
- [Cu=63.5, Ag=108, 1F = 96,500C]

Data

$$I = 0.2 \text{ A}$$

$$m_{\text{Ag}} = 0.780 \text{ g}$$

$$m_{\text{Cu}} = ?$$

$$Q = ?$$

$$t = ?$$

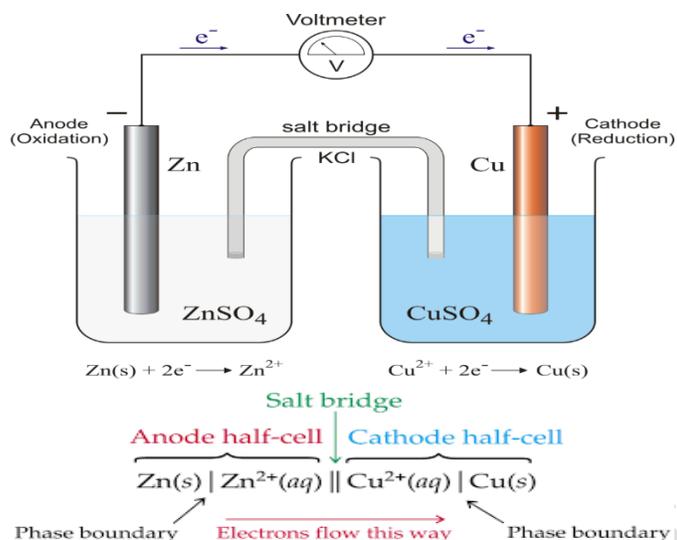
(Practice on questions that does not require formulas to be solved)

Assignment 9

1. Explain rusting/corrosion as an electrochemical process
2. Explain cathodic protection

Electrochemical cell

An electrochemical cell is a device that can generate electrical energy from the chemical reactions occurring in it.



Cell Notation

The anode half-cell is described first; the cathode half-cell follows

A single vertical line (|) is drawn between two chemical species that are in different phases but in physical contact with each other.

A double vertical line (//) represents a salt bridge or porous membrane separating the individual half-cells.

The phase of each chemical (s, l, g, aq) is shown in parentheses.

Standard Electrode Potential, E^\ominus

The standard electrode potential of a metal ions/metal system is the potential difference set up between the metal and a one-molar solution of its ions at 25 °C (arbitrarily taking the standard electrode potential of the hydrogen ions/hydrogen gas system as zero volt).

Electrode potentials vary from one metal ions/ metal system to another. The electrode potential of a given system depends on

- the overall energy change,
- the concentration of ions in the solution, and
- the temperature.

Calculating the Electromotive Force (emf) of a cell, E_{cell}

$$E_{\text{cell}}^{\ominus} = E_{\text{red}}^{\ominus} - E_{\text{oxid}}^{\ominus}$$

$$E_{\text{cell}}^{\ominus} = E_{\text{cathode}}^{\ominus} - E_{\text{anode}}^{\ominus}$$

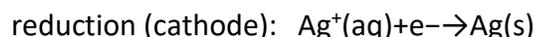
Example

Calculate the standard cell potential of a voltaic cell that uses the Ag/Ag⁺ and Sn/Sn²⁺ half-cell reactions. Write the balanced equation for the overall cell reaction that occurs. Identify the anode and the cathode ($E^{\circ}_{\text{Ag}}=+0.80\text{V}$ and $E^{\circ}_{\text{Sn}}=-0.14\text{V}$)

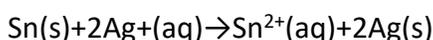
Solution:

$$E^{\circ}_{\text{cell}}=?\text{V}$$

- The silver half-cell will undergo reduction. The tin half-cell will undergo oxidation.



- overall equation:**



- $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{red}} - E^{\circ}_{\text{oxid}}$
 $= +0.80 - (-0.14\text{V}) = +0.94\text{V}$

- Think about your result.*

The standard cell potential is positive, so the reaction is spontaneous as written. Tin is oxidized at the anode, while silver ion is reduced at the cathode. Note that the voltage for the silver ion reduction is not doubled even though the reduction half-reaction had to be doubled to balance the overall redox equation.

(Solve more problems from prep 50)

Types of Electrochemical cell

- Primary cell; are basically use-and-throw galvanic cells. The electrochemical reactions that take place in these cells are irreversible in nature. Hence, the reactants are consumed for the generation of electrical energy and the cell stops producing an electric current once the reactants are completely depleted. Examples are Daniel cell, leclanche cell/dry cell.
- Secondary cells; (also known as rechargeable batteries) are electrochemical cells in which the cell has a reversible reaction, i.e. the cell can function as a Galvanic cell as well as an Electrolytic cell. E.g. lead acid accumulator

Difference between an Electrolytic Cell and an Electrochemical cell

Galvanic Cell / Voltaic Cell	Electrolytic Cell
Chemical energy is transformed into electrical energy in these electrochemical cells.	Electrical energy is transformed into chemical energy in these cells.

The redox reactions that take place in these cells are spontaneous in nature.	An input of energy is required for the redox reactions to proceed in these cells, i.e. the reactions are non-spontaneous.
In these electrochemical cells, the anode is negatively charged and the cathode is positively charged.	These cells feature a positively charged anode and a negatively charged cathode.
The electrons originate from the species that undergoes oxidation.	Electrons originate from an external source (such as a battery).

Assignment 10

1. With the aid of a well labelled diagram explain the working principle of a Leclanche cell
2. Explain with the aid of equations the discharging and recharging process of a lead acid accumulator

Week 11: Radioactivity

This is the spontaneous emission of radiation by a radioactive substance

Types of Radiation

1. **Alpha (α) particle, ${}^4_2\text{He}$**

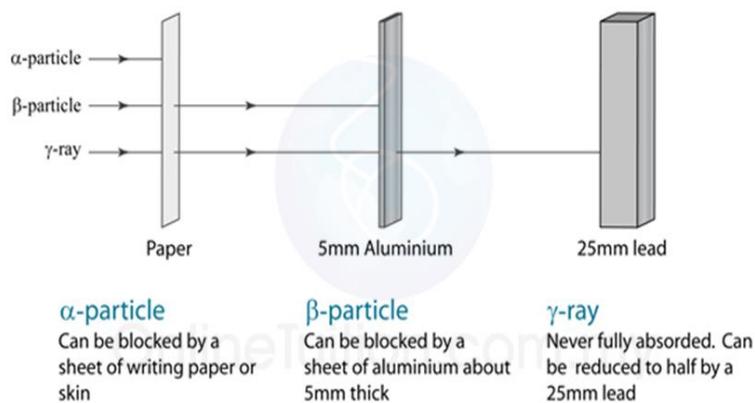
 - It is a helium nucleus
 - positively charged
 - low penetrating power

2. **Beta (β) particle, ${}^0_{-1}\text{e}$**

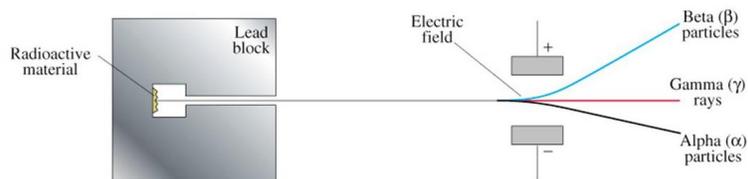
 - It is an electron
 - negatively charged
 - higher penetrating power
 - relatively lower ionizing power

3. **Gamma-ray (γ)**

 - It is an electromagnetic wave
 - no charge
 - highest penetrating power
 - least ionizing power.



Separating Alpha, Beta, and Gamma Particles



Alpha and beta particles are deflected in opposite directions- alpha particles toward the negative plate and beta particles toward the positive plate. Gamma rays are undeflected.

Detection of Radiation

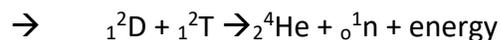
1. The Geiger Muller counters
2. Scintillation detector
3. Diffusion cloud chamber

Alpha, Beta, and Gamma decay

Decay Type	Radiation Emitted	Generic Equation	Model
Alpha decay	${}^4_2\alpha$	${}^A_ZX \longrightarrow {}^{A-4}_{Z-2}X' + {}^4_2\alpha$	 <p>Parent → Daughter + Alpha Particle</p>
Beta decay	${}^0_{-1}\beta$	${}^A_ZX \longrightarrow {}^A_{Z+1}X' + {}^0_{-1}\beta$	 <p>Parent → Daughter + Beta Particle</p>
Gamma emission	${}^0_0\gamma$	${}^A_ZX^* \xrightarrow{\text{Relaxation}} {}^A_ZX' + {}^0_0\gamma$	 <p>Parent (excited nuclear state) → Daughter + Gamma ray</p>

Nuclear Fusion

this is defined as the process in which two or more light nuclei fuse or combine to form a heavier nucleus with the release of radiation and energy



Nuclear fission

a process in which the nucleus of a heavy element is split into two nuclei of nearly equal mass with a release of energy and radiation



Student: Ensure you practice past exam questions on decay

Teacher: explain the working principle of a nuclear power plant

Comparing Chemical and Nuclear reactions

Chemical Reactions	Nuclear Reactions
1. Atoms are rearranged by the breaking and forming of chemical bonds.	1. Elements (or isotopes of the same elements) are converted from one to another.
2. Only electrons in atomic or molecular orbitals are involved in the breaking and forming of bonds.	2. Protons, neutrons, electrons, and other elementary particles may be involved.
3. Reactions are accompanied by absorption or release of relatively small amounts of energy.	3. Reactions are accompanied by absorption or release of tremendous amounts of energy.
4. Rates of reaction are influenced by temperature, pressure, concentration, and catalysts.	4. Rates of reaction normally are not affected by temperature, pressure, and catalysts.

Assignment 11

1. With the aid of appropriate equations, differentiate between artificial transmutation and nuclear fusion and fission
2. Outline 5 uses of radioactivity and radioisotopes

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