This is a Class 1 exam

Lebanon Evangelical School for Boys and Girls

Loueizeh - Baabda

SE01 Exam

10NOV2025

Do NOT write in this box. Your MARK will be placed here.

Subject :	Chemistry	
Class :	11S1 and 11S2	
Teacher :	Mrs. Mirella Freige	
Duration :	1 hour	
You are to answer the questions to this exam on : Lined paper		
Your name :		*Time taken to complete exam*
Your Class :		
Exam Room :		
How many sheets, including cover sheet :		

- Do not open this paper until the teacher in charge instructs you to do so.
- The most important thing to remember in taking an exam is NEVER PANIC.
- If you are caught cheating in this exam, you will be awarded zero for ALL exams that are taken on this day.

Exercise 1 Redox reaction (6 points)

Given: Molar masses in g/mol: M($AgNO_3$) =170; M(Ag)= 108 and M(Cu)=64.

A copper wire (Cu) with a mass of 3.2 g is dipped into 100 mL of silver nitrate solution ($Ag^+ + NO_3^-$) of molar concentration C.

A redox reaction took place : $Cu_{(s)} + 2Ag^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$

A- Preparation of silver nitrate solution

The silver nitrate solution is prepared by dissolving 3.4 g of solid silver nitrate ($AgNO_3$) in 100 mL of water.

1- Choose from the following list, the materials needed to prepare the above solution.

- Volumetric pipettes : 5, 10 and 20 mL.
- Volumetric flasks : 100, 250 and 500 mL.
- Digital balance, watch glass and spatula.
- Funnel and distilled water.

2- Show that the concentration of the prepared solution is C = 0.2 mol/L.

B- Redox reaction

1- Describe your observations a few minutes after the reaction started.

2- Indicate the two redox couples involved and classify them.

3- Determine the limiting reactant.

4- Calculate the concentration of copper ions Cu^{2+} obtained.

Exercise 2 Reaction between aluminum and iron II ions (6 points)

Given: Molar masses in g/mol: M(Al) = 27; M(Fe)= 56.

An aluminum strip (Al) with a mass of 5.4 g is completely dipped into 50 mL of iron II sulfate solution ($Fe^{2+} + SO_4^{2-}$) of concentration 0.2 mol/L. After a few minutes, the aluminum strip becomes coated with a deposit of iron metal Fe.

The equation of the redox reaction is $\mathbf{2Al + 3Fe^{2+} \rightarrow 2Al^{3+} + 3Fe}$

1- Write the oxidation and the reduction half equations.

2- Indicate the two redox couples involved and classify them.

3- Show that the initial mixture is non-stoichiometric. Deduce the limiting reactant.

4- Calculate the mass of iron formed.

5- Calculate the mass of aluminum remaining at the end of the reaction.

Exercise 3 Qualitative study of a redox reaction (8 points)

1- Which of the following statements is incorrect?

a) The reducing agent undergoes an oxidation.

b) The oxidizing agent is called oxidant.

c) The oxidant gains electrons.

d) The reductant gains electrons.

2- Given the unbalanced equation: $Hg + Au^{3+} \rightarrow Hg^{2+} + Au$

a) Indicate the redox couples involved.

b) Balance the above equation.

c) Identify the strongest oxidizing agent and the strongest reducing agent.

3- Consider the electrochemical classification of the following redox couples:



a) Predict the possible redox reactions and write the overall balanced equation for each reaction.

b) Gold ions Au^{3+} can oxidize silver Ag and produce a gold deposit Au.

Place the Au^{3+} / Au couple in the above classification.

SE01/20252026/Chemistry/11S1S2

Exercise 1 Redox reaction (6 points)

A- Preparation of silver nitrate solution:

1- The materials needed:

- Volumetric flask: 100 mL
- Digital balance, watch glass and spatula
- Funnel and distilled water. (1)

2- $m_{AgNO_3(s)} = 3.4\ g$ $\quad M = 170\ g/mol$.

$V_{solution} = 100\ mL$.

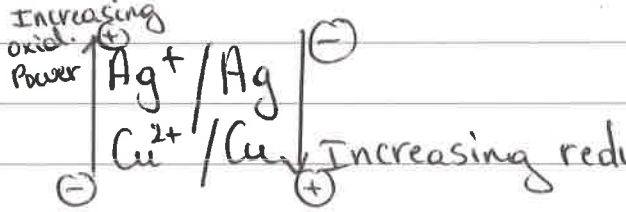
$$C_{solution} = \frac{n_{solute}}{V_{solution}} = \frac{m_{solute}}{M_{solute} \times V_{solution}} = \frac{3.4}{170 \times 100 \times 10^{-3}}$$

$C_{solution} = 0.2\ mol/L$. (1)

B- Redox reaction

1- . The mass of Cu wire decreases.
. The solution becomes blue.
. A silvery solid is obtained. (3/4)

2- The two redox couples are: Ag^+/Ag and Cu^{2+}/Cu.



(3/4)

3- Initial number of moles for each reactant:

$n_{0,Cu} = \frac{m}{M} = \frac{3.2}{64} = 5 \times 10^{-2}\ mol$ (1/4)

$n_{0,Ag^+} = C \times V = 0.2 \times 100 \times 10^{-3} = 2 \times 10^{-2}\ mol$ (1/4)

$R_{Cu} = \frac{n_{o\,Cu}}{coeff} = \frac{5\times10^{-2}}{1} = 5\times10^{-2}$ (1/4)

$R_{Ag^+} = \frac{n_{o\,Ag^+}}{coeff} = \frac{2\times10^{-2}}{2} = 10^{-2}$ (1/4)

$R_{Ag^+} < R_{Cu}$ (1/4) $\Rightarrow$ Ag^+: Limiting reactant. (1/4)

4- $Cu_{(s)} + 2Ag^+_{(aq)} \longrightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$

According to stoichiometry:

$$\frac{n_{Ag^+ reacted}}{2} = \frac{n_{Cu^{2+} obtained}}{1}$$

$$\Rightarrow \frac{n_{o\,Ag^+}}{2} = n_{Cu^{2+}\,obtained}$$

$$\Rightarrow n_{Cu^{2+}} = \frac{2\times10^{-2}}{2} = 10^{-2}\ mol.$$ (1/2)

$$[Cu^{2+}] = \frac{n_{Cu^{2+}}}{V_{solution}} = \frac{10^{-2}}{100\times10^{-3}} = 10^{-1}\ mol\times L^{-1}.$$ (1/2)

Exercise 2 Reaction between Al and Fe^{2+} ions (6 points)

$m_{Al} = 5.4\ g$ $\quad V_{(Fe^{2+}+SO_4^{2-})} = 50\ mL$

$M_{Al} = 27\ g/mol.$ $\quad C = 0.2\ mol/L.$

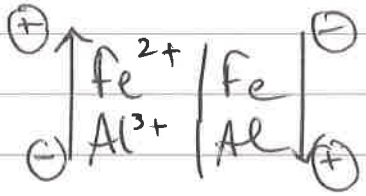
The equation is $2Al + 3Fe^{2+} \longrightarrow 2Al^{3+} + 3Fe$

1- $Al_{(s)} \longrightarrow 3\bar{e} + Al^{3+}$: oxidation (1/2)

$Fe^{2+} + 2\bar{e} \longrightarrow Fe_{(s)}$: reduction (1/2)

2- Couples: Al^{3+}/Al and Fe^{2+}/Fe.

The redox reaction takes place between the stronger reductant (Al) and the stronger oxidant (Fe^{2+})



(3/4)

3- $n_{o\,Al} = \frac{m}{M} = \frac{5.4}{27} = 0.2$ mol.

$n_{o\,Fe^{2+}} = C \times V = 0.2 \times 50 \times 10^{-3} = 10^{-2}$ mol.

$R_{Al} = \frac{n_{o\,Al}}{\text{coeff.}} = \frac{0.2}{2} = 10^{-1}$

$R_{Fe^{2+}} = \frac{n_o}{\text{coeff}} = \frac{10^{-2}}{3} = 0.033 \times 10^{-1}$

$R_{Al} \neq R_{Fe^{2+}} \longrightarrow$ initial mixture is non-stoichiometric

$R_{Fe^{2+}} < R_{Al} \Rightarrow Fe^{2+}$: Limiting reactant. ($1\frac{1}{2}$)

4- Stoichiometry: $\frac{n_{Al\ reacted}}{2} = \frac{n_{Fe^{2+}\ reacted}}{3} = \frac{n_{Fe\ obtained}}{3}$

$\Rightarrow n_{Fe\ obtained} = n_{o\,Fe^{2+}} = 10^{-2}$ mol.

Mass of Fe $= n \times M_{Fe} = 10^{-2} \times 56 = 0.56$ g. ($1\frac{1}{2}$)

5- $n_{Al\ reacted} = \frac{2 \times n_{Fe^{2+}\ reacted}}{3} = \frac{2 \times 10^{-2}}{3} = 0.66 \times 10^{-2}$ mol.

$\Rightarrow n_{Al\ remaining} = n_{o\,Al} - n_{Al\ reacted}$

$= 0.2 - 0.66 \times 10^{-2}$

$= 0.193$ mol.

$m_{Al\ remaining} = n \times M_{Al}$

$= 0.193 \times 27 = 5.211$ g. ($1\frac{1}{4}$)

Exercise 3 (8 points)

1- d- (1)

2- $Hg + Au^{3+} \longrightarrow Hg^{2+} + Au$.

a- Redox couples: Au^{3+}/Au and Hg^{2+}/Hg (1)

b- Balanced equation:

$3Hg + 2Au^{3+} \longrightarrow 3Hg^{2+} + 2Au$ (1)

c- Au^{3+} oxidant is stronger than Hg^{2+} and Hg reductant is stronger than Au. (1)

3-a) The redox reaction takes place between the strongest oxidant and the strongest reductant.

. Ag^+ is stronger than Fe^{2+}

. Ag^+ is stronger than Zn^{2+}

$\Rightarrow$ Ag^+ can react with Fe : $2Ag^+ + Fe \longrightarrow 2Ag + Fe^{2+}$ (1)

Ag^+ can react with Zn : $2Ag^+ + Zn \longrightarrow 2Ag + Zn^{2+}$ (1)

Fe^{2+} oxidant is stronger than Zn^{2+} (1)

$\rightarrow$ Fe^{2+} can react with Zn : $Zn + Fe^{2+} \longrightarrow Zn^{2+} + Fe$.

b) Au^{3+} ions can react with Ag because Au^{3+} oxidant is stronger than Ag^+ $\rightarrow$ Au^{3+}/Au is positioned above Ag^+/Ag.

$\uparrow$ Au^{3+}/Au

Ag^+/Ag $\downarrow$ (1)