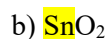
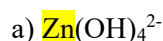


**Exercise 1.** Identify the oxidation number of the element labeled in yellow in each of the following compounds:



**Exercise 2.** Identify the oxidation number of the element labeled in yellow in each of the following ions:



*Hint: Don't forget to take into consideration the ionic state.*

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**Exercise 3.** When atomic nickel (Ni) is added to  $\text{CuCl}_2$  (aq),  $\text{Ni}^{2+}$  ions and atomic copper (Cu) form. When iron is added to  $\text{NiCl}_2$  (aq),  $\text{Fe}^{2+}$  ions and atomic nickel (Ni) form.

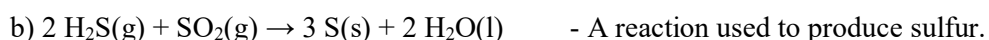
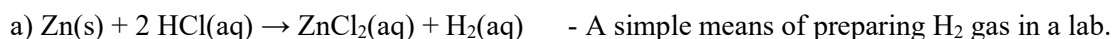
a) What will happen if atomic form of iron (Fe) is added to  $\text{CuCl}_2$  (aq)? Explain.

b) What are Fe and  $\text{Cu}^{2+}$  in the context of redox properties?

c) If you wanted to restore  $\text{Fe}^{2+}$  back to Fe, could you achieve that using Cu or Ni as electron donors?

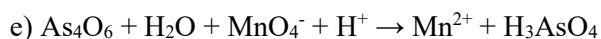
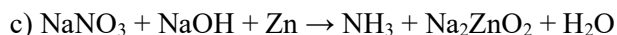
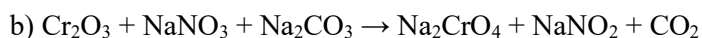
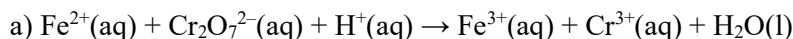
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**Exercise 4.** Identify the oxidant and reductant in each of the following reactions:

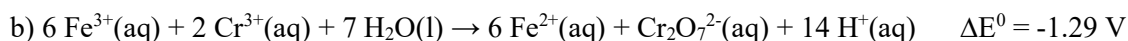


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**Exercise 5.** Determine the stoichiometric coefficients of the following reactions:



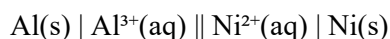
**Exercise 6.** Calculate the standard reaction Gibbs free energy for the following cell reactions:



$F = 96485 \text{ C/mol}$

---

**Exercise 7.** Consider the galvanic cell below:



- Write the corresponding half-reactions for each electrode.
- Assign which electrode will serve as cathode and which will serve as anode in this cell.
- What will be the direction of this reaction (i.e., which species will undergo oxidation and which species will be reduced)? Considering this, write a balanced chemical equation for this redox reaction.
- Calculate the standard electromotive force of the cell ( $\Delta E^0$ ).
- Calculate the  $\Delta G_r^0$  and determine if the reaction will favor reactants or products.
- Determine the equilibrium constant. What is that number suggesting about the reaction propensity under standard conditions.

The given values are:

$$E^0_{\text{Ni}^{2+}/\text{Ni}} = -0.257 \text{ V};$$

$$E^0_{\text{Al}^{3+}/\text{Al}} = -1.676 \text{ V}$$

$$F = 96485 \text{ C/mol}$$

$$T = 298 \text{ K}$$

$$R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$$

---

**Exercise 8.** Consider the Daniell cell made of a zinc plate immersed in a solution of  $\text{ZnSO}_4$  whose activity of  $\text{Zn}^{2+}$  is 1 (concentration = 1 mol/L), and a copper plate immersed in a solution of  $\text{CuSO}_4$ ; the two electrodes are connected by a salt bridge. When the current is zero (high-resistance voltmeter used), the potential difference measured between the terminals of the cell is  $\Delta E = 1.02 \text{ V}$ .

- Is the system operating under standard conditions? Explain your rationale.
- What is the potential of the  $\text{Zn}^{2+}/\text{Zn}$  and  $\text{Cu}^{2+}/\text{Cu}$  electrodes?
- Calculate the concentration of  $\text{Cu}^{2+}$  in the second compartment (assuming that it perfectly matches  $\text{Cu}^{2+}$  activity).

The given values:

$$E^\circ(\text{Zn}^{2+}/\text{Zn}) = -0.762 \text{ V} ;$$

$$E^\circ(\text{Cu}^{2+}/\text{Cu}) = 0.342 \text{ V}$$

$$T = 298 \text{ K, atmospheric pressure.}$$

$$F = 96485 \text{ C/mol}$$

$$R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$$

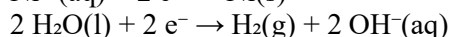
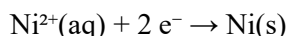
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**Exercise 9.** Consider the thermodynamic cell at  $25^\circ\text{C}$  composed of two electrodes connected by a salt bridge: the first is a copper strip immersed in 1 liter of a  $\text{Cu}(\text{NO}_3)_2$  solution in which the activity of  $\text{Cu}^{2+}$  is 1 (concentration = 1 mol/L), and the second electrode is also a copper strip immersed in 1 liter of a  $\text{Cu}(\text{NO}_3)_2$  solution in which the activity of  $\text{Cu}^{2+}$  is  $5 \times 10^{-4}$  (concentration =  $5 \times 10^{-4}$  mol/L).

- In this cell system, what is the cathode and what is the anode?
- Determine the potential difference ( $\Delta E$ ) between the two electrodes when the current is zero (i.e., high resistance voltmeter used).
- What are the concentrations of  $\text{Cu}^{2+}$  in the two compartments (electrodes) when the cell is fully discharged (dead)?
- What is then the change in mass of the two copper strips when the cell is fully discharged?

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**Exercise 10.** Ni electrodes can often cause the reduction of  $\text{H}_2\text{O}$  and formation of  $\text{H}_2$  gas during normal operation. Let's assume that you have such electrode connected into an electrolytic cell. A current of 15 A is used to deposit solid Ni from an aqueous solution of  $\text{NiSO}_4$ . Nickel metal and hydrogen gas form simultaneously at the same electrode according to the reactions below. The system uses 60% of the current to produce Ni and 40% to produce  $\text{H}_2$ .



- What is the mass of Ni deposited during 1 hour?
- What is the volume of  $\text{H}_2$  released at  $25^\circ\text{C}$  and the pressure of 1 bar during the same time? Assume that  $\text{H}_2$  is a perfect gas .

Given values:

$$F = 96485 \text{ C/mol}$$

$$R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$$

$$M(\text{Ni}) = 58.693 \text{ g mol}^{-1}$$

$$M(\text{H}) = 1.0078 \text{ g mol}^{-1}$$

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Quick Answers:

1.

a) +4; b) +4; c) -2; d) +5; e) +1; f) 0;

2.

a) +2; b) +2; c) +6; d) +4; e) +1;

3.

a)  $\text{Fe}^{2+}$  and Cu will form. b) Fe is the reducing agent;  $\text{Cu}^{2+}$  is the oxidizing agent. c) No.

4.

a) Oxidant:  $\text{H}^+$  in HCl; Reductant: Zn;

b) Oxidant:  $\text{SO}_2$ ; Reductant:  $\text{H}_2\text{S}$ ;

c) Oxidant:  $\text{B}_2\text{O}_3$ ; Reductant: Mg;

d) Oxidant:  $\text{N}_2\text{O}_4$ ; Reductant:  $\text{N}_2\text{H}_4$ ;

5.

a)  $6 \text{Fe}^{2+}(\text{aq}) + 1 \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14 \text{H}^+(\text{aq}) \rightarrow 6 \text{Fe}^{3+}(\text{aq}) + 2 \text{Cr}^{3+}(\text{aq}) + 7 \text{H}_2\text{O}(\text{l})$

b)  $1 \text{Cr}_2\text{O}_3 + 3 \text{NaNO}_3 + 2 \text{Na}_2\text{CO}_3 \rightarrow 2 \text{Na}_2\text{CrO}_4 + 3 \text{NaNO}_2 + 2 \text{CO}_2$

c)  $1 \text{NaNO}_3 + 7 \text{NaOH} + 4 \text{Zn} \rightarrow 1 \text{NH}_3 + 4 \text{Na}_2\text{ZnO}_2 + 2 \text{H}_2\text{O}$

d)  $3 \text{NaHSO}_4 + 8 \text{Al} + 3 \text{NaOH} \rightarrow 3 \text{Na}_2\text{S} + 4 \text{Al}_2\text{O}_3 + 3 \text{H}_2\text{O}$

e)  $5 \text{As}_4\text{O}_6 + 18 \text{H}_2\text{O} + 8 \text{MnO}_4^- + 24 \text{H}^+ \rightarrow 8 \text{Mn}^{2+} + 20 \text{H}_3\text{AsO}_4$

6.

a)  $\Delta G_r^0 = -208'407 \text{ J/mol}$

b)  $\Delta G_r^0 = 746'793 \text{ J/mol}$

7.

a)  $\text{Ni}^{2+}(\text{aq}) + 2 \text{e}^- \rightarrow \text{Ni}(\text{s})$

$\text{Al}^{3+}(\text{aq}) + 3 \text{e}^- \rightarrow \text{Al}(\text{s})$

b) Cathode: Nickel

Anode: Aluminium

c)  $3 \text{Ni}^{2+}(\text{aq}) + 2 \text{Al}(\text{s}) \rightarrow 3 \text{Ni}(\text{s}) + 2 \text{Al}^{3+}(\text{aq})$

d)  $\Delta E^0 = 1.419 \text{ V}$

e)  $\Delta G_r^0 = -821'473 \text{ J/mol}$ ; Strongly favouring products

f)  $K = 9.9 \cdot 10^{143}$ ; The reaction will STRONGLY favor the products.

8.

a) No, since  $\Delta E^0$  is not equal to  $\Delta E$ .

b)  $E(\text{Zn}^{2+}/\text{Zn}) = -0.762$ ;  $E(\text{Cu}^{2+}/\text{Cu}) = 0.258 \text{ V}$ ;

c)  $[\text{Cu}^{2+}] = 1.44 \cdot 10^{-3} \text{ mol/L}$

9.

a) Compartment 1 ( $[\text{Cu}^{2+}]_1 = 1 \text{ mol/L}$ ) – cathode; Compartment 2 ( $[\text{Cu}^{2+}]_2 = 5 \cdot 10^{-4} \text{ mol/L}$ ) – anode.

b)  $\Delta E^0 = 0.098 \text{ V}$

c)  $[\text{Cu}^{2+}]_1 = [\text{Cu}^{2+}]_2 = 0.50025 \text{ mol/L}$

d) Cathode gained 31.76 g; Anode lost 31.76 g;

10.

a)  $m(\text{Ni}) = 9.85 \text{ g}$

b)  $V(\text{H}_2) = 2.77 \text{ L}$