

(EXAM GEOLOGY2)

Exercise 1 (5 points)

Definitions (1 point each)

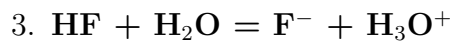
1. **Oxidation:** Loss of electrons by a chemical species → increase in oxidation number. 1 pts
 2. **Reduction:** Gain of electrons by a chemical species → decrease in oxidation number. 1 pts
 3. **Oxidizing agent (oxidant):** Species that accepts electrons → it gets reduced. 1 pts
 4. **Reducing agent (reductant):** Species that donates electrons → it gets oxidized. 1 pts
 5. **Rules for balancing redox equations** (half-reaction method): 1 pts
 - Write the oxidation and reduction half-reactions.
 - Balance atoms other than H and O.
 - Balance O by adding H₂O.
 - Balance H by adding H⁺ (acidic) or OH⁻ (basic).
 - Balance charge by adding electrons.
 - Multiply half-reactions to equalize electrons.
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Exercise 2 (8 points)

Part 1 – Conjugate acid/base pairs (4 points)

1. $\text{NH}_3 + \text{H}_2\text{O} = \text{NH}_4^+ + \text{OH}^-$
 - Acid: H₂O
 - Base: NH₃ 1.25pts
 - Conjugate acid: NH₄⁺
 - Conjugate base: OH⁻
2. $\text{HCO}_3^- + \text{H}_2\text{O} = \text{CO}_3^{2-} + \text{H}_3\text{O}^+$
 - Acid: H₂O

- Base: HCO_3^-
- Conjugate acid: H_3O^+ 1.25pts
- Conjugate base: CO_3^{2-}



- Acid: HF
- Base: H_2O 1.5pts
- Conjugate acid: H_3O^+
- Conjugate base: F^-

Part 2 – pH calculation (2 points)

Given: NaOH solution, $C = 6.4 \times 10^{-4}$ M
 Strong base $\rightarrow [\text{OH}^-] = C$

$$\text{pOH} = -\log(6.4 \times 10^{-4}) = 3.19 \quad 1\text{pts}$$

$$\text{pH} = 14 - \text{pOH} = 14 - 3.19 = \boxed{10.81} \quad 1\text{pts}$$

Part 3 – pH classification (2 points)

Given: $\text{pH}_1 = 2.5$, $\text{pH}_2 = 7.0$, $\text{pH}_3 = 11.2$, $\text{pH}_4 = 4.8$, $\text{pH}_5 = 9.3$

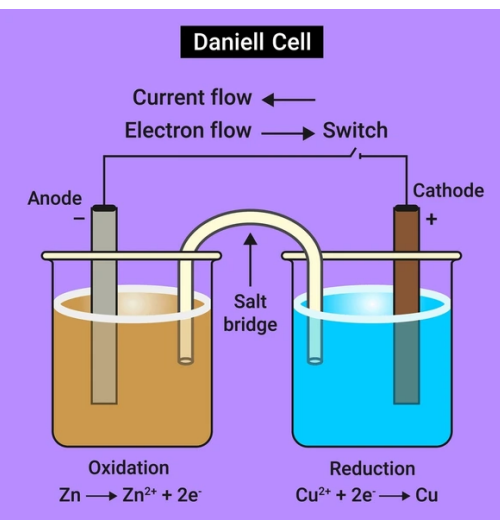
1. Classification:

- Acidic: pH_1 , pH_4
- Neutral: pH_2 1pts
- Basic: pH_3 , pH_5

2. Order from most acidic to most basic:

$$\text{pH}_1(2.5) < \text{pH}_4(4.8) < \text{pH}_2(7.0) < \text{pH}_5(9.3) < \text{pH}_3(11.2) \quad 1\text{pts}$$

Exercise 3 (7 points)



1.5pts

1. Electrode potentials (2 points)

Using the Nernst equation at 25°C:

$$E = E^\circ + \frac{0.059}{n} \log[M^{n+}] \quad \text{0.5pts}$$

For Ni electrode (anode):

$$E_{\text{Ni}} = -0.257 + \frac{0.059}{2} \log(0.2)$$

$$\log(0.2) = -0.6990$$

$$\frac{0.059}{2} = 0.0295 \quad \text{1pts}$$

$$E_{\text{Ni}} = -0.257 + 0.0295 \times (-0.6990) = -0.257 - 0.0206 = -0.2776 \text{ V}$$

For Cu electrode (cathode):

$$E_{\text{Cu}} = 0.340 + \frac{0.059}{2} \log(0.2)$$

$$E_{\text{Cu}} = 0.340 + 0.0295 \times (-0.6990) = 0.340 - 0.0206 = 0.3194 \text{ V} \quad \text{1pts}$$

2. EMF of the cell (1 point)

$$\text{EMF} = E_{\text{cathode}} - E_{\text{anode}} = E_{\text{Cu}} - E_{\text{Ni}}$$

$$\text{EMF} = 0.3194 - (-0.2776) = 0.3194 + 0.2776 = \boxed{0.5970 \text{ V}} \quad \text{1pts}$$

3. Equilibrium constant K (2 points)

Using the empirical Nernst equation at 25°C:

$$\log K = \frac{n \times E_{\text{cell}}^\circ}{0.059}$$

$$\log K = \frac{2 \times 0.597}{0.059} = \frac{1.194}{0.059} \approx 20.24 \quad \text{1pts}$$

$$K = 10^{20.24} = 10^{0.24} \times 10^{20} \approx \boxed{1.74 \times 10^{20}}$$

4. Type of electrodes (1 point)

Both are **metal electrodes of the first kind** (metal in direct contact with its own ions):

1pts

- Ni electrode: Ni | Ni²⁺ (0.2 M)
- Cu electrode: Cu | Cu²⁺ (0.2 M)