

Exercise walk-through

9.9

Cell Potential and Free Energy

AP Chemistry

Free energy and equilibrium

Free energy connects to the equilibrium constant:

Standard and nonstandard

$$\Delta G^\circ = -RT \ln K$$

$$\Delta G = \Delta G^\circ + RT \ln Q$$

$\Delta G^\circ < 0 \Rightarrow K > 1$. At **equilibrium**, $\Delta G = 0$ and $Q = K$.

Coupled reactions

Add the ΔG° values. An unfavorable step (+20) coupled to -50:

$$+20 + (-50) = -30 \text{ kJ}$$

In cells, the “downhill” partner is ATP breakdown.

Exercises 2.1-2.3

2.1 Write E_{cell}° in terms of the electrode potentials. [1]

2.2 For cathode $+0.80$ V and anode -0.44 V, find E_{cell}° . [2]

2.3 A cell has a positive E_{cell}° . Is it galvanic or electrolytic? [1]

Exercises 2.1-2.3

2.1 Write E_{cell}° in terms of the electrode potentials. [1]

2.1

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

2.2 For cathode +0.80 V and anode -0.44 V, find E_{cell}° . [2]

2.2

- $0.80 - (-0.44) = +1.24 \text{ V}$

2.3 A cell has a positive E_{cell}° . Is it galvanic or electrolytic? [1]

2.3

- Galvanic

Exercise 3.1

3.1 A galvanic (favorable) cell has an E°_{cell} that is

[1]

☐

A positive

B negative

C zero

D undefined

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[1]

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A positive

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3.1 A

- positive

Exercise 3.2 (a)

3.2 A cell has $E_{\text{cathode}}^{\circ} = +0.34 \text{ V}$ and $E_{\text{anode}}^{\circ} = -0.76 \text{ V}$, with $n = 2$.

(a) Find E_{cell}° .

[2]

Exercise 3.2 (a)

3.2 A cell has $E_{\text{cathode}}^{\circ} = +0.34 \text{ V}$ and $E_{\text{anode}}^{\circ} = -0.76 \text{ V}$, with $n = 2$.

(a) Find E_{cell}° .

[2]

(a)

- $E_{\text{cell}}^{\circ} = 0.34 - (-0.76) = +1.10 \text{ V}$

Exercise 3.2 (b)

3.2 A cell has $E_{\text{cathode}}^{\circ} = +0.34 \text{ V}$ and $E_{\text{anode}}^{\circ} = -0.76 \text{ V}$, with $n = 2$.

(b) Find ΔG° (use $F = 96\,500 \text{ C/mol}$).

[3]

Exercise 3.2 (b)

3.2 A cell has $E_{\text{cathode}}^{\circ} = +0.34 \text{ V}$ and $E_{\text{anode}}^{\circ} = -0.76 \text{ V}$, with $n = 2$.

(b) Find ΔG° (use $F = 96\,500 \text{ C/mol}$).

[3]

(b)

- $\Delta G^{\circ} = -(2)(96\,500)(1.10) = -2.12 \times 10^5 \text{ J} = -212 \text{ kJ}$

Exercise 3.3

3.3 Explain why a positive E_{cell}° corresponds to a negative ΔG° .

[2]

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3.3

- $\Delta G^{\circ} = -nFE^{\circ}$; with n and F positive, a positive E° makes ΔG° negative – a favorable reaction