

Solutions to Chm139 Practice Exam

Part A: Multiple Choice Questions

1. When a 1.00 g sample of a mixture of NaCl and AlCl₃ is treated with excess AgNO₃ solution, 0.02142 moles of solid AgCl are obtained. What is the percentage by mass of NaCl in the mixture? (NaCl, 58.44 g·mol⁻¹; AlCl₃, 133.33 g·mol⁻¹)
- A) less than 10%
 B) 20%
 C) 25%
 D) 31%
 E) 42%

Answer: D**Solution:** $4\text{AgNO}_3 + \text{NaCl} + \text{AlCl}_3 \rightarrow 4\text{AgCl(s)} + \text{NaNO}_3 + \text{Al(NO}_3)_3$

$$0.02141 \text{ moles AgCl} \times \frac{1 \text{ mole NaCl}}{4 \text{ moles AgCl}} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mole NaCl}} = 0.313 \text{ g NaCl}$$

$$\% \text{ mass NaCl} = \frac{0.313 \text{ g}}{1.00 \text{ g}} \times 100\% = 31\%$$

2. Which of the following sets of quantum numbers is not possible?

	n	l	m _l	m _s
A.	1	0	0	+1/2
B.	4	0	0	+1/2
C.	3	3	-3	-1/2
D.	2	1	1	-1/2
E.	2	0	0	+1/2

Answer: C**Solution:** $l = 0, 1, \dots (n-1)$, therefore if $n = 3$, then l cannot equal 3.

3. Only one of the following molecules or ions is **not** of VSEPR type AX₃. Which one is it?

- A) CO₃²⁻
 B) SO₃
 C) NO₃⁻
 D) KrF₃⁺
 E) BF₃

Answer: D

Solution: KrF₃⁺ has lone pairs on the central atom giving AX₃E₂.

4. The minimum energy for photoemission of electrons from a potassium surface is 3.69 x 10⁻¹⁹J. What is the kinetic energy of electrons emitted from a potassium surface when it is irradiated by UV light at 300nm.

- A) 7.60 x 10⁻²⁰J
 B) 3.69 x 10⁻¹⁹J
 C) 2.93 x 10⁻¹⁹J
 D) 6.62 x 10⁻¹⁹J
 E) 1.03 x 10⁻¹⁸J

Answer: C

Solution: $E = hc/\lambda = 6.626 \times 10^{-34} \text{ Js}^{-1} (3.0 \times 10^8 \text{ ms}^{-1}) / (300 \times 10^{-9} \text{ m})$
 $= 6.626 \times 10^{-19} \text{ J}$

Kinetic energy = $6.626 \times 10^{-19} \text{ J} - 3.69 \times 10^{-19} \text{ J} = 2.93 \times 10^{-19} \text{ J}$

5. By law, microwave ovens are allowed to operate only at a fixed frequency of 2450 MHz. To cook a typical hot dog requires 27.4 kJ of heat. How many microwave photons must a typical hot dog absorb to be cooked in a microwave oven?

- A) 2.17 x 10²⁷
 B) 1.69 x 10²⁸
 C) 2.17 x 10²⁴
 D) 1.43 x 10⁸
 E) 1.68 x 10²⁵

Answer: B

Solution: Energy of one photon = $h\nu = (6.626 \times 10^{-34} \text{ Js})(2450 \times 10^6 \text{ s}^{-1}) = 1.623 \times 10^{-24} \text{ J}$

$$27.4 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ photon}}{1.623 \times 10^{-24} \text{ J}} = 1.69 \times 10^{28}$$

6. For the reaction $A + B \rightarrow \text{products}$, the following data were obtained.

$[A]_0$ (in mol L ⁻¹)	$[B]_0$ (in mol L ⁻¹)	Initial Rate (in mol L ⁻¹ min ⁻¹)
0.40	0.40	160.0
0.20	0.40	80.0
0.20	0.10	5.00
0.40	0.20	40.0

What is the rate law for this reaction?

- A) Rate = $k[A][B]$
 B) Rate = $k[A]^2[B]$
 C) Rate = $k[A][B]^2$
 D) Rate = $k[A]^2[B]^4$
 E) Rate = $k[A]^2[B]^2$

Answer: C

Solution: Let the m be the order with respect to $[A]_0$ and let n be the order with respect to $[B]_0$. Let the rate be: $\text{Rate} = k[A]^m[B]^n$.

To solve for m:

(cancel out all like terms) then

$$\frac{160}{80} = \frac{[0.40]^m}{[0.20]^m}$$

$$2 = \left[\frac{0.40}{0.20} \right]^m$$

$$2 = 2^m, \text{ therefore } : m = 1$$

To solve for n:

$$\frac{\text{exp1}}{\text{exp4}} = \frac{\text{Rate1}}{\text{Rate4}} = \frac{k[A]^m[B]^n}{k[A]^m[B]^n} \Rightarrow \frac{160}{40} = \frac{k[0.40]^m[0.40]^n}{k[0.40]^m[0.20]^n} \quad (\text{cancel out all like terms})$$

then:

$$= \frac{160}{40} = \frac{[0.40]^n}{[0.20]^n}$$

$$4 = \left[\frac{0.40}{0.20} \right]^n$$

$$4 = 2^n$$

$$n = 2$$

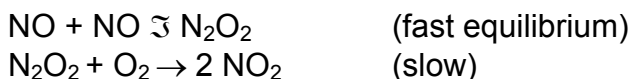
Therefore, when $m=1$ and $n=2$, the rate is: **Rate = $k[A][B]^2$**

7. The half-life of a radioactive isotope is ten days. How many days does it take for the isotope to decay to one-eighth of its original activity? (Radioactivity decay is a first-order process).
- A) 20 days
B) 30 days
C) 32 days
D) 48 days
E) 80 days

Answer: B

Solution: In one half-life it will decay to one-half its original activity, two half-lives to one-quarter, and three half-lives to one-eighth. Therefore it would take 30 days (or three half-lives) to decay to one eighth of the original activity.

8. The reaction of NO with O₂ to form NO₂ contributes to the formation of urban smog. The following mechanism has been proposed:



What is the rate law derived from this mechanism?

- A) Rate = constant \times [NO]² [O₂]
B) Rate = constant \times [NO]²
C) Rate = constant \times [NO] [O₂]²
D) Rate = constant \times [NO] [O₂]
E) Rate = constant \times [O₂]²

Answer: A

Solution: N₂O₂ is an intermediate, therefore cannot be present in the rate law.

Rate is found from the slower of the two reactions:

$$\text{Rate} = k[\text{N}_2\text{O}_2][\text{O}_2], \text{ where } k[\text{N}_2\text{O}_2] = k_{-1}[\text{NO}]^2$$

$$\text{Therefore rate} = \text{constant}[\text{NO}]^2[\text{O}_2]$$

9. The rate of a certain reaction is about 3.0 times faster at 50°C than at 40°C. Which of the following is the best estimate of the activation energy for the reaction?
- A) 10 J mol⁻¹
 - B) 50 J mol⁻¹
 - C) 90 J mol⁻¹
 - D) 200 J mol⁻¹
 - E) 400 J mol⁻¹

Answer: C

Solution: rate = $e^{-E_a/RT}$, then $\Delta\text{rate} = e^{-E_a/R\Delta T}$
 $3 = e^{-E_a/[(8.314\text{J/molK})(-10\text{K})]}$

$$\ln 3 = \frac{E_a}{83.14\text{J/mol}} = 91\text{ J/mol}$$

10. Consider the reaction
- $$3\text{Fe(s)} + 4\text{H}_2\text{O(g)} \rightleftharpoons 4\text{H}_2\text{(g)} + \text{Fe}_3\text{O}_4\text{(s)}$$

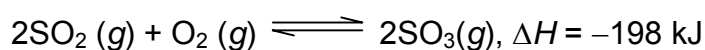
If the total pressure is increased suddenly by reducing the volume,

- A) more H₂O (g) is produced
- B) no change occurs
- C) more H₂(g) is produced
- D) more Fe(s) is produced

Answer: B

Solution: No change will occur as there are equal moles of gases on either side of the equation. The change in pressure will have no effect on the solid species, they are not part of the K_{eq} constant.

11. The following reaction has reached a state of dynamic equilibrium, in the presence of a catalyst.



Which one of the following changes will decrease the concentration of SO_3 in the equilibrium mixture?

- A) increasing the temperature
- B) removing the catalyst
- C) decreasing the volume of the reaction vessel
- D) adding some more O_2
- E) none of these

Answer: A

Solution: Removing the catalyst has no effect and C & D will shift the equilibrium to the right.

12. If 0.00300 mol NaOH and 0.00300 mol $\text{Ba}(\text{OH})_2$ are completely dissolved in water to make 1.00 L of solution at 25°C , what is the pOH of the solution?

- A) 2.05
- B) 2.22
- C) 7.00
- D) 11.8
- E) 12.0

Answer: A

Solution: Because both NaOH and $\text{Ba}(\text{OH})_2$ are strong bases, we can assume that they will completely dissociate, therefore:

$$0.00300 \text{ mol NaOH} \times 1 \text{ mol OH}^- / 1 \text{ mol NaOH} = 0.003 \text{ mol OH}^-$$

$$0.00300 \text{ mol Ba}(\text{OH})_2 \times 2 \text{ mol OH}^- / 1 \text{ mol Ba}(\text{OH})_2 = 0.006 \text{ mol OH}^-$$

$$\text{Total moles OH}^- = 0.009 \text{ moles} = 0.009 \text{ M}$$

$$\text{pOH} = -\log([\text{OH}^-]) = -\log(0.009) = 2.05$$

13. Recall the equation $\text{pH} = \text{p}K_{\text{HIn}} + \log\left(\frac{[\text{In}^-]}{[\text{HIn}]}\right)$. For the acid-base indicator HIn, $K_a = 1 \times 10^{-6}$. At $\text{pH} = 8.0$, what is the ratio of acid form of the indicator to its conjugate base form?
- A) 200/1
B) 1/200
C) 50/1
D) 1/100
E) 1/1

Answer: D

Solution: For this equilibrium $[\text{H}^+] = [\text{In}^-] = 10^{-8} = 1 \times 10^{-8}$

$$K_a = 1 \times 10^{-6} = \frac{[\text{H}^+][\text{In}^-]}{[\text{HIn}]} = \frac{(1 \times 10^{-8})^2}{[\text{HIn}]}$$

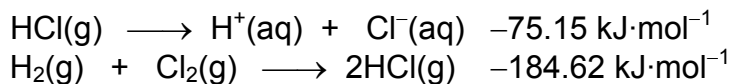
$$[\text{HIn}] = 1 \times 10^{-10}, \text{ therefore } [\text{HIn}]/[\text{In}^-] = 1 \times 10^{-10}/1 \times 10^{-8} = 0.01 = 1/100$$

14. Calculate the work needed to make room for products in the combustion of 1 mole of C(s) at 273 K and 1 atm. (1 L.atm = 101 J)
- A) -2.26 kJ
B) no work is needed
C) -6.79 kJ
D) -4.52 kJ

Answer: B

Solution: $\text{C(s)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$, no work is needed as there is one mole of gas on both the product and the reaction side of the equation.

15. Given the following thermochemical data at 1 atm:

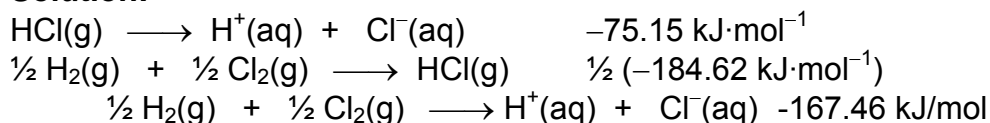


Determine the standard enthalpy of formation of $\text{Cl}^-(\text{aq})$

- A) $+17.16 \text{ kJ}\cdot\text{mol}^{-1}$
 B) $+109.47 \text{ kJ}\cdot\text{mol}^{-1}$
 C) $-167.46 \text{ kJ}\cdot\text{mol}^{-1}$
 D) $+34.32 \text{ kJ}\cdot\text{mol}^{-1}$

Answer: C

Solution:

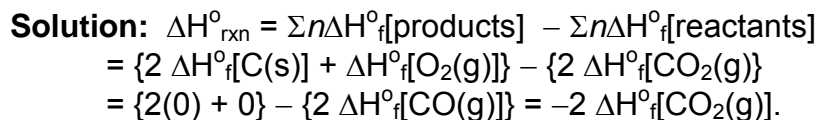


16. What is the standard reaction enthalpy for the reaction below?



- A) $-\Delta H_f^\circ[\text{CO}_2(\text{g})]$
 B) $+\Delta H_f^\circ[\text{CO}_2(\text{g})]$
 C) $-2 \Delta H_f^\circ[\text{CO}_2(\text{g})]$
 D) $+2 \Delta H_f^\circ[\text{CO}_2(\text{g})]$
 E) $+2 \Delta H_f^\circ[\text{C(s)}] + \Delta H_f^\circ[\text{O}_2(\text{g})]$

Answer: C



C(s) and $\text{O}_2(\text{g})$ are the most stable elements at 25°C and 1 atm. Hence, $\Delta H_f^\circ[\text{C(s)}] = 0$ and $\Delta H_f^\circ[\text{O}_2(\text{g})] = 0$.

17. When 5.00g H_2SO_4 (formula weight, 98.09 g/mol) is dissolved in water in a calorimeter, the temperature decreases by 0.718°C . If the heat capacity of the calorimeter (and its contents) is $6.83 \text{ kJ } ^\circ\text{C}^{-1}$, what is the molar heat of solution of H_2SO_4 ?
- A) -481 kJ/mol
B) -10.3 kJ/mol
C) -96.2 kJ/mol
D) -0.526 kJ/mol
E) -0.981 kJ/mol

Answer: C

Solution:

$$\Delta T = 0.718^\circ\text{C}$$

$$C_p = 6.83 \text{ kJ } ^\circ\text{C}^{-1}$$

$$\text{moles } \text{H}_2\text{SO}_4 = 5.00\text{g} (1\text{mol}/ 98.09\text{g}) = 0.051 \text{ moles}$$

$$\Delta H^\circ = -\Delta T C_p/\text{moles} = -96.2 \text{ kJ/mol}$$

18. The reaction $\text{CO}(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$

is extremely important in efforts to reduce poisonous emissions from automobile exhausts. At constant T and P, which of the following is correct for this reaction?

- A) $w = 0$
B) $\Delta H = \Delta U$
C) ΔH is less than ΔU
D) ΔH is greater than ΔU
E) ΔH is independent of the physical states of the reactants and products.

Answer: C

Solution: At constant P, $q_p = \Delta H \Rightarrow \Delta U = \Delta H + w$

$$\Delta n(\text{g}) = 1 - 1 - \frac{1}{2} = -\frac{1}{2}. \quad w = -P(\Delta V) = -[\Delta n(\text{g})]RT = -(-1/2)RT = +ve$$

Hence, $\Delta U > \Delta H$ or ΔH is less than ΔU .

19. "Spontaneous change is always accompanied by an increase in entropy". The entropy change referred to in this statement is:

- A) ΔS_{system}
 B) $\Delta S_{\text{universe}}$
 C) $\Delta S_{\text{surroundings}}$
 D) $\Delta S_{\text{zero Kelvin temperature}}$

Answer: B

Solution: Spontaneous change is always accompanied by an increase in entropy. This is the second law of thermodynamics.

$$\Delta S_{\text{universe}} = \Delta S_{\text{surroundings}} + \Delta S_{\text{system}} > 0. \text{ Hence, the entropy change referred to } \Delta S_{\text{universe}}.$$

20. The vapor pressure of formic acid, $\text{HCOOH}_{(l)}$, at 25°C is 11.4 Torr.



The change in free energy for the reaction, ΔG_r , at 25°C is:

- A) $+ (0.008314)(298) \ln (1) = 0$
 B) $+ (0.008314)(298) \ln (11.4) = 6.03\text{kJ/mol}$
 C) $- (0.008314)(298) \ln (11.4) = -6.03\text{kJ/mol}$
 D) $- (0.008314)(298) \ln (0.015) = 10.4\text{kJ/mol}$
 E) $+ (0.008314)(298) \ln (0.015) = -10.4\text{kJ/mol}$

Answer: A

Solution: Given: $\text{HCOOH}_{(l)} \rightarrow \text{HCOOH}_{(g)} \quad (11.4 \text{ Torr})$

To Find: ΔG_r 25°C (298K)

For phase transition at equilibrium, $K_{\text{eq}} = Q$, $\Delta G = 0$

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

$$\Delta G = -RT \ln K_{\text{eq}} + RT \ln Q = RT \ln (Q / K_{\text{eq}}) = RT \ln (1) \quad (\text{for phase transition})$$

$$= (0.008314 \text{ kJ/mol/K}) (298\text{K}) \ln (1)$$

$$= 0$$

$$\begin{aligned} \text{However } \Delta G^\circ_r &= -RT \ln K_{\text{eq}} \\ &= - (0.008314 \text{ kJ/mol/K}) (298\text{K}) \ln (11.4/760) \\ &= + 10.4 \text{ kJ/mol} \end{aligned}$$

21. Which of the following reactions has the largest positive molar entropy change?

- A) $4\text{NH}_3(\text{g}) + \text{N}_2(\text{g}) \rightarrow 3\text{N}_2\text{H}_4(\text{l})$
 B) $\text{C}(\text{s}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{CO}(\text{g}) + \text{H}_2(\text{g})$
 C) $\text{NH}_4\text{NO}_3(\text{s}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{NO}_3^-(\text{aq})$
 D) $\text{Na}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NaOH}(\text{s}) + \frac{1}{2} \text{H}_2(\text{g})$
 E) $\text{N}_2(\text{g}) (1\text{atm}, 25^\circ\text{C}) \rightarrow \text{N}_2(\text{g}) (10\text{atm}, 25^\circ\text{C})$

Answer: B

Solution: Factors affecting entropy change:

$$S(\text{g}) > S(\text{l}) > S(\text{s})$$

ΔS_r° is positive when $\Delta n(\text{g})$ is positive

ΔS_r° is positive when ΔV increases (i.e. P decreases)

- | | |
|--|--|
| A) $4\text{NH}_3(\text{g}) + \text{N}_2(\text{g}) \rightarrow 3\text{N}_2\text{H}_4(\text{l})$ | $\Delta n(\text{g}) = 0 - 4 - 1 = -5$ ($\Delta S = -\text{ve}$) |
| B) $\text{C}(\text{s}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{CO}(\text{g}) + \text{H}_2(\text{g})$ | $\Delta n(\text{g}) = 1 + 1 - 1 = +1$ ($\Delta S = +\text{ve}$) |
| C) $\text{NH}_4\text{NO}_3(\text{s}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{NO}_3^-(\text{aq})$ | $S(\text{s}) < S(\text{aq})$ ($\Delta S = +\text{ve}$, but less +ve) |
| D) $\text{Na}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NaOH}(\text{s}) + \frac{1}{2} \text{H}_2(\text{g})$ | $\Delta n(\text{g}) = +\frac{1}{2}$ ($\Delta S = +\text{ve}$) |
| E) $\text{N}_2(\text{g}) (1\text{atm}, 25^\circ\text{C}) \rightarrow \text{N}_2(\text{g}) (10\text{atm}, 25^\circ\text{C})$ | P increases, V decreases, ($\Delta S \downarrow$) |

22. When sulfur, $\text{S}_8(\text{s})$, burns in oxygen to produce $\text{SO}_2(\text{g})$ under standard conditions it generates 2375 kJ of heat per mole of S_8 burned. What is the standard molar enthalpy of formation of $\text{SO}_2(\text{g})$?

- A) 297 kJ/mol
 B) 2375 kJ/mol
 C) -2375 kJ/mol
 D) 594 kJ/mol
 E) -297 kJ/mol

Answer: A

Solution: $\frac{1}{8} \text{S}_8(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g}) \quad \Delta H_f^\circ = ?$

$$\Delta H_f^\circ = \Delta H_{\text{comb}}^\circ / 8 = 2375 \text{kJ/mol} \times 1/8 = 297 \text{kJ/mol}$$

23. The N_2 , F_2 , and N-F bond enthalpies are 944, 158, and 272 $\text{kJ}\cdot\text{mol}^{-1}$ respectively. Calculate the enthalpy of formation of $\text{NF}_3(\text{g})$.
- A) +365
 - B) -1525
 - C) -816
 - D) -107

Answer: D

Solution: $\text{NF}_3(\text{g})$ has 3 N-F bonds forming and the breaking on one N_2 bond and two F_2 bonds.

$$\Delta H_f = \text{bond breaking} - \text{bond forming} = [0.5(944\text{kJ/mol}) + 2/3(158\text{kJ/mol})] - 3(272)$$

kJ/mol

$$\Delta H_f = -107 \text{ kJ/mol}$$

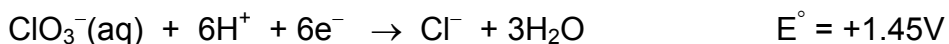
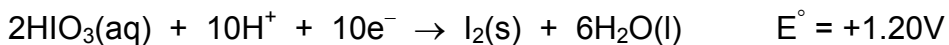
24. Which one of the following statements is **true**?

- A) ΔH is always equal to q .
- B) ΔH is always equal to ΔU .
- C) ΔH equals to ΔU for a reaction which is carried out at constant volume.
- D) ΔH equals q for a reaction which is carried out at constant pressure.
- E) ΔH for an endothermic reaction is negative.

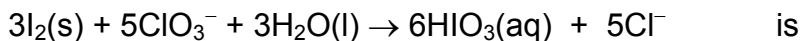
Answer: D

Solution: ΔH only equals q if the reaction is carried out at constant pressure. ΔH does not always equal ΔU . ΔH for an endothermic reaction is positive.

25. Given the following standard reduction potentials



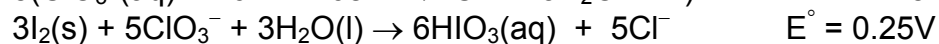
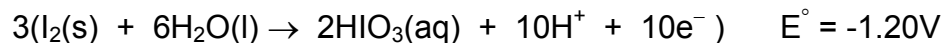
The standard cell potential for the reaction



- A) +0.25V
- B) -2.48V
- C) +2.65V
- D) -3.65V
- E) +3.65V

Answer: A

Solution:



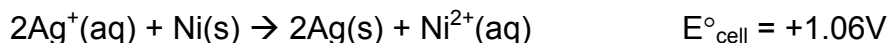
26. In a redox reaction, the reducing agent:

- A) gains electrons and is reduced
- B) gains electrons and is oxidized
- C) contains an element that undergoes a decrease in oxidation number
- D) loses electrons and is reduced
- E) loses electrons and is oxidized

Answer: E

Solution: A reducing agent causes another species to be reduced; thus it is itself oxidized. Oxidized species loses electrons.

27. Consider the reaction



If E° for the Ag reduction half-reaction is +0.80V, the standard reduction potential for

$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ni}(\text{s})$ at 25°C
in V, must be...

- A) -0.28
- B) -0.26
- C) +0.54
- D) -0.54
- E) +0.26

Answer: B

Solution: In the reaction given above, Ag^+ is being reduced to Ag while Ni is being oxidized to Ni^{2+} . The overall E°_{cell} for the reaction is calculated as

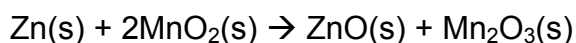
$$E^\circ_{\text{cell}} = E_{\text{reduction}} - E_{\text{oxidation}}$$

The reduction half of the reaction is given as +0.80V. In order to calculate the E value for the oxidation half, rearrange the E°_{cell} equation to solve for $E_{\text{oxidation}}$.

$$E_{\text{oxidation}} = E^\circ_{\text{cell}} - E_{\text{reduction}} = (+1.06\text{V}) - (+0.80\text{V}) = +0.26\text{V}.$$

The $E_{\text{oxidation}}$ value is the opposite sign from the $E_{\text{reduction}}$ value, so in order to solve for the $E_{\text{reduction}}$ value as asked, the sign on the $E_{\text{oxidation}}$ value must be reversed. Therefore, $E_{\text{reduction}}$ for $\text{Ni}^{2+} = -0.26\text{V}$

28. If an alkaline dry cell provides a steady current of 0.40A for ten hours, through the reaction



the mass in g of zinc metal consumed is...

- A) 4.9
- B) 2.4
- C) 9.8
- D) 20
- E) 1.2

Answer: A

Solution: Calculate the coulombs of charge that are produced by the reaction from the given variables. Then determine how many moles of electrons are found in the coulombs calculated (using Faraday's constant). Then determine how many moles of Zn are needed. Convert the moles of Zn to the mass. Let n equal the moles and let m equal the mass.

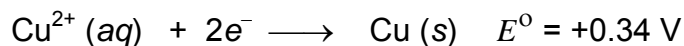
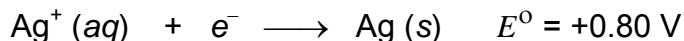
$$\text{Coulombs} = \text{Amperes} \times \text{Time} = (0.40\text{A})(36,000\text{sec}) = 14,400\text{C}$$

$$n_{e^-} = \frac{14,400\text{C}}{96,490\text{C/mol}} = 0.149\text{mol } e^-$$

$$n_{\text{Zn}} = 0.149\text{mol } e^- \times \frac{1\text{mol Zn}}{2\text{mol } e^-} = 0.0745\text{mol Zn}$$

$$m_{\text{Zn}} = 0.0745\text{mol} \times 65.39\text{g/mol} = 4.9\text{g}$$

29. A strip of copper is placed in a 1 M solution of copper nitrate and a strip of silver is placed in a 1M solution of silver nitrate. The two metal strips are connected to a voltmeter by wires and a salt bridge connects the solutions. The following standard reduction potentials apply:



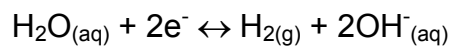
When the voltmeter is removed and the two electrodes are connected by a wire, which of the following does not take place?

- A) Electrons flow in the external circuit from the copper electrode to the silver electrode.
- B) The silver electrode increases in mass as the cell operates.
- C) There is a net general movement of silver ions through the salt bridge to the copper half-cell.
- D) Negative ions pass through the salt bridge from the silver half-cell to the copper half-cell.
- E) Some positive copper ions pass through the salt bridge from the copper half-cell to the silver half-cell.

Answer: C

Solution: In this cell, Ag^+ is being reduced to $\text{Ag}(\text{s})$ at the cathode and $\text{Cu}(\text{s})$ is being oxidized to Cu^{2+} at the anode. Current flows from the anode to the cathode. There will be a net movement of copper ions through the salt bridge to the silver half-cell to balance some of the charges.

30. What is the volume of the H₂ gas at 25°C and 1.00 atm that will collect at the cathode when an aqueous solution of Na₂SO₄ is electrolyzed for 2.00 hours with a current of 10.0 A?



- A) 9.13L
- B) 0.373L
- C) 18.26L
- D) 4.55L
- E) 0.766L

Answer: A

Solution: $\frac{It}{96485, \text{Cmol}^{-1}} = \text{moles of electrons}$

$$\frac{(10.0\text{A})(7200\text{sec})}{96485\text{C/mol}} = 0.75 \text{ moles electrons. Since there are two moles e}^- \text{ per mole H}_2(\text{g}):$$

There will be 0.373 moles H₂(g). We can then use PV = nRT

$$(1\text{atm})V = (0.373\text{mol})(0.0821\text{atmL/molK})(298\text{K})$$

$$V = 9.13\text{L}$$

Part B: Written Answer Questions

1. Iron crystallizes in one of the cubic lattice arrangements. The edge length of the unit cell is 287 pm. The density of iron is 7.86 g cm^{-3} at 298K and its molar mass is $55.847 \text{ g mol}^{-1}$. What type of lattice does iron form? Show your reasoning.

Answer: There are two atoms per unit cell. The unit cell is a body-centred cubic.

Solution:

$$\left(\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} \right) \left(\frac{55.847 \text{ g}}{1 \text{ mol}} \right) \left(\frac{1 \text{ cm}^3}{7.86 \text{ g}} \right) \left(\frac{1 \text{ pm}^3}{1 \times 10^{-30} \text{ cm}^3} \right) = \frac{1.18 \times 10^7 \text{ pm}^3}{\text{atom}}$$

The volume of the unit cell is $(287 \text{ pm})^3 = 2.36 \times 10^7 \text{ pm}^3$

$$\text{Atoms per unit cell} = (2.36 \times 10^7 \text{ pm}^3) \left(\frac{1.18 \times 10^7 \text{ pm}^3}{\text{atom}} \right) = 2 \text{ atoms per unit cell.}$$

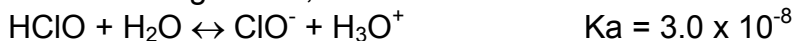
Therefore, the cell is body centred.

2. If you have a buffer solution made up of 0.25mol hypochlorous acid (HOCl), $K_a = 3.0 \times 10^{-8}$, and 0.40mol sodium hypochlorite (NaOCl) in 1.0L of water, calculate the change in pH of this solution after 0.12mol of NaOH are added. Assume no volume change.

Solution:

$$[\text{HOCl}] = 0.25 \text{ M}, \quad [\text{ClO}^-] = 0.40 \text{ M}$$

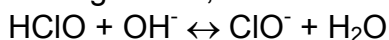
Before adding NaOH,



$$3.0 \times 10^{-8} = [\text{H}_3\text{O}^+](0.40 \text{ M}) / (0.25 \text{ M})$$

$$[\text{H}_3\text{O}^+] = 1.9 \times 10^{-8} \text{ M} \quad \text{pH} = 7.73$$

Adding NaOH,



$$\text{Mol ClO}^- = 0.40 \text{ mol} + 0.12 \text{ mol} = 0.52 \text{ mol}$$

$$\text{Mol HClO} = 0.25 \text{ mol} - 0.12 \text{ mol} = 0.13 \text{ mol}$$

$$3.0 \times 10^{-8} = [\text{H}_3\text{O}^+](0.52 \text{ M}) / (0.13 \text{ M})$$

$$[\text{H}_3\text{O}^+] = 7.5 \times 10^{-9} \text{ M} \quad \text{pH} = 8.12$$

$$\text{Change in pH} = 8.12 - 7.73 = 0.39$$

3. Exposure of metallic iron to moist air results in the formation of rust, one component of which is iron(III) oxide, Fe₂O₃. Metallic iron can be recovered from iron(III) oxide by heating it in the presence of solid graphitic carbon:



For the substances involved in this reaction we have the following data at 25°C:

Substance	ΔH_f° (kJ mol ⁻¹)	S° (J mol ⁻¹ K ⁻¹)
Fe ₂ O ₃ (s)	-824.2	87.4
C(graphite)		5.7
Fe(s)		27.3
CO ₂ (g)	-393.5	213.7

Assume ideal behavior.

- (a) Calculate ΔH° for the above reaction at 25°C.

Answer:

$$\Delta H^\circ = 467.9 \text{ kJ}$$

$$\begin{aligned} \Delta H^\circ &= \sum n_p \Delta H_f^\circ(\text{prod}) - \sum n_r \Delta H_f^\circ(\text{react}) \\ &= (4 \times 0 + 3(-393.5)) - (2(-824.2) + 3(0)) \\ &= +467.9 \text{ kJ} \end{aligned}$$

- (b) Calculate ΔS° for the above reaction at 25°C.

Answer:

$$\Delta S^\circ = 558.4 \text{ JK}^{-1}$$

$$\begin{aligned} \Delta S^\circ &= \sum n_p S^\circ(\text{prod}) - \sum n_r S^\circ(\text{react}) \\ &= (4(27.3) + 3(213.7)) - (2(87.4) + 3(5.7)) \\ &= 558.4 \text{ JK}^{-1} \end{aligned}$$

- (c) Account for the algebraic sign of ΔS° for the above reaction.

Answer:

Entropy (disorder) increases (ie. ΔS° is positive) because 5 moles of solid are connected to 4 moles of solid and 3 moles of gas.

- (d) Calculate ΔG° for the above reaction at 25 °C.

Answer:

$$301.5 \text{ kJ}$$

Solution

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = 467.9 - 298 \times 558.4/1000 = 301.5 \text{ kJ}$$

(e) Calculate the equilibrium constant, K , for the above reaction at 25°C.

Answer:

$$K = 1.4 \times 10^{-53}$$

$$\ln K = \frac{-\Delta G^0}{RT} = -\frac{301,500}{8.314 \times 298} = -121.7$$

$$K = 1.4 \times 10^{-53}$$

(f) Determine the temperature above which this reaction would become spontaneous under standard state conditions.

Answer:

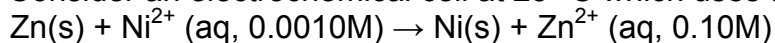
838K

Solution

ie. We require $\Delta G^0 \leq 0$

$$\text{ie } \Delta G^0 = 0 \text{ when } T = \frac{\Delta H^0}{\Delta S^0} = \frac{467,900 \text{ J}}{558.4 \text{ JK}^{-1}} = 838 \text{ K}$$

4. Consider an electrochemical cell at 25 °C which uses the chemical reaction:



to which the following standard reduction potentials apply

$$\left\{ \begin{array}{lll} \text{Ni}^{2+}(\text{aq}) + 2\text{e}^- & \text{Ni(s)} & \varepsilon_{298}^0 = -0.23 \text{ volt} \\ \text{Zn}^{2+}(\text{aq}) + 2\text{e}^- & \text{Zn(s)} & \varepsilon_{298}^0 = -0.76 \text{ volt} \end{array} \right\} \varepsilon_{\text{cell}}^0 = 0.53 \text{ volt}$$

(a) Calculate the value for the reaction quotient, Q , for the cell reaction, as written above.

Answer:

$$Q = 100$$

Solution:

$$Q = \frac{a_{\text{Ni}} a_{\text{Zn}^{2+}}}{a_{\text{Zn}} a_{\text{Ni}^{2+}}} = \frac{(1)(0.1)}{(1)(0.001)} = 100$$

(b) Determine the cell potential at 25°C.

Answer:

Cell potential = 0.47 volt

Solution

$$\varepsilon_{\text{cell}} = \varepsilon_{\text{cell}}^0 - \frac{RT}{nF} \ln Q = 0.53 - \frac{8.314 \times 298}{2 \times 96,500} \ln 100 = 0.47 \text{ volt}$$

c) Calculate ΔG for the cell at 25 °C.

Answer:

-90.71kJ

$$\Delta G^\circ = -nF\varepsilon_{cell} = -2 \times 96,500 \times 0.47 = -90.71 \text{ kJ}$$

- d) Specify the composition of the electrode at which reduction occurs during the spontaneous operation of the cell.

Answer: Ni