ACS GENERAL CHEMISTRY 2 PRACTICE EXAM (SAMPLE)

STUDY GUIDE



SAMPLE STUDY GUIDE WITH LIMITED QUESTIONS

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Introduction

Preparing for a certification exam can feel overwhelming, but with the right tools, it becomes an opportunity to build confidence, sharpen your skills, and move one step closer to your goals. At Examzify, we believe that effective exam preparation isn't just about memorization, it's about understanding the material, identifying knowledge gaps, and building the test-taking strategies that lead to success.

This guide was designed to help you do exactly that.

Whether you're preparing for a licensing exam, professional certification, or entry-level qualification, this book offers structured practice to reinforce key concepts. You'll find a wide range of multiple-choice questions, each followed by clear explanations to help you understand not just the right answer, but why it's correct.

The content in this guide is based on real-world exam objectives and aligned with the types of questions and topics commonly found on official tests. It's ideal for learners who want to:

- Practice answering questions under realistic conditions,
- Improve accuracy and speed,
- Review explanations to strengthen weak areas, and
- Approach the exam with greater confidence.

We recommend using this book not as a stand-alone study tool, but alongside other resources like flashcards, textbooks, or hands-on training. For best results, we recommend working through each question, reflecting on the explanation provided, and revisiting the topics that challenge you most.

Remember: successful test preparation isn't about getting every question right the first time, it's about learning from your mistakes and improving over time. Stay focused, trust the process, and know that every page you turn brings you closer to success.

Let's begin.

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How to Use This Guide

This guide is designed to help you study more effectively and approach your exam with confidence. Whether you're reviewing for the first time or doing a final refresh, here's how to get the most out of your Examzify study guide:

1. Start with a Diagnostic Review

Skim through the questions to get a sense of what you know and what you need to focus on. Your goal is to identify knowledge gaps early.

2. Study in Short, Focused Sessions

Break your study time into manageable blocks (e.g. 30 – 45 minutes). Review a handful of questions, reflect on the explanations.

3. Learn from the Explanations

After answering a question, always read the explanation, even if you got it right. It reinforces key points, corrects misunderstandings, and teaches subtle distinctions between similar answers.

4. Track Your Progress

Use bookmarks or notes (if reading digitally) to mark difficult questions. Revisit these regularly and track improvements over time.

5. Simulate the Real Exam

Once you're comfortable, try taking a full set of questions without pausing. Set a timer and simulate test-day conditions to build confidence and time management skills.

6. Repeat and Review

Don't just study once, repetition builds retention. Re-attempt questions after a few days and revisit explanations to reinforce learning. Pair this guide with other Examzify tools like flashcards, and digital practice tests to strengthen your preparation across formats.

There's no single right way to study, but consistent, thoughtful effort always wins. Use this guide flexibly, adapt the tips above to fit your pace and learning style. You've got this!

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Questions

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1. Which statement best describes colligative properties?

A. They depend on solute identity

B. They depend on solvent dielectric constant

C. They depend on molality and number of particles: ΔTb = iK_b m; ΔTf = iK_f m

D. They depend only on temperature

2. How does a catalyst affect a reaction's activation energy and rate?

A. It lowers Ea and increases rate without altering equilibrium

B. It raises Ea and decreases rate

C. It lowers Ea but changes equilibrium constant

D. It increases Ea and increases rate

3. Which acid is the weakest in aqueous solution?

A. acetic acid (Ka = 1.8 × 10^-5)

B. formic acid (Ka = 1.8 × 10^-4)

C. hydrocyanic acid (Ka = 6.2 × 10^-10)

D. nitrous acid (Ka = 4.5 × 10^-4)

4. During electrolysis of molten sodium iodide, what is the half-reaction occurring at the anode?

A. I- -> I+ + e-

B. Na -> Na+ + e-

C. 2 I- -> I2 + 2 e-

D. I2 + 2 e- -> 2 I-

5. In acidic solution, the half-reactions during electrolysis of water are:

A. Cathode: 2 H2O + 2 e− H2 + 2 OH−; Anode: 2 H2O O2 + 4 H+ + 4 e−

B. Cathode: 2 H+ + 2 e− H2; Anode: H2O 1/2 O2 + 2 H+ + 2 e−

C. Cathode: 4 H+ + 4 e− 2 H2; Anode: O2 + 4 e− 2 H2O

D. Cathode: H2O + e− 1/2 H2 + OH−; Anode: H2O H+ + e−

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6. What is the overall balanced equation for the electrolysis of water in acidic solution?

A. $2\ H_2O \rightarrow 2\ H_2 + O_2$

B. $2\ H_2O \rightarrow H_2 + O_2$

C. $H_2O \rightarrow H_2 + 1/2\ O_2$

D. $4\ H_2O \rightarrow 2\ H_2 + O_2$

7. Calculate the standard Gibbs free energy change at 298 K for the reaction $FeO_4\ (s) + 4\ CO\ (g) \rightarrow 3\ Fe\ (s) + 4\ CO_2\ (g)$.

A. -14.8 kJ/mol

B. -757 kJ/mol

C. 757 kJ/mol

D. -757 kJ/mol

8. In beta decay, what happens to the atomic number Z?

A. Increases by 1

B. Increases by 2

C. Decreases by 1

D. Remains the same

9. Standard reduction potentials can be used to judge spontaneity of a redox reaction by:

A. If $E°cell < 0$ for the cell as written, the reaction is spontaneous

B. If $E°cell = 0$ for the cell as written, the reaction is spontaneous

C. If $E°cell > 0$ for the cell as written, the reaction is spontaneous

D. If E°cell is undefined, the reaction is spontaneous

10. Given the standard reduction potentials at 25 degrees C, what is the standard cell potential of a $Ni\ (s)\ |\ Ni^{2+}\ (aq)\ ||\ Ag^{+}\ (aq)\ |\ Ag\ (s)$ galvanic cell?

A. 0.56 V

B. 1.02 V

C. 1.35 V

D. 1.81 V

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Answers

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1. C
2. A
3. C
4. C
5. B
6. A
7. A
8. A
9. C
10. A

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Explanations

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1. Which statement best describes colligative properties?

A. They depend on solute identity

B. They depend on solvent dielectric constant

C. They depend on molality and number of particles: ΔTb = iK_b m; ΔTf = iK_f m

D. They depend only on temperature

Colligative properties depend on the number of dissolved particles, not on what those particles are. In dilute solutions, the boiling-point elevation and freezing-point depression are quantified by ΔTb = i Kb m and ΔTf = i Kf m, where m is molality and i is the van't Hoff factor that accounts for how many particles the solute produces in solution. This means the magnitude of these effects scales with how many particles are present and with concentration, rather than the specific identity of the solute. Therefore, the correct description emphasizes molality and particle count, as shown by that relationship. The other ideas don't fit: the identity of the solute doesn't determine colligative effects, the solvent's dielectric constant isn't the defining factor, and these properties are not determined solely by temperature.

2. How does a catalyst affect a reaction's activation energy and rate?

A. It lowers Ea and increases rate without altering equilibrium

B. It raises Ea and decreases rate

C. It lowers Ea but changes equilibrium constant

D. It increases Ea and increases rate

A catalyst provides an alternative reaction pathway with a lower activation energy, so more molecules have enough energy to react and the rate increases. It speeds up both forward and reverse reactions equally, meaning the equilibrium position and the equilibrium constant stay the same. So the catalyst lowers Ea, raising the rate, without altering the equilibrium. Statements that suggest raising Ea or changing the equilibrium are not consistent with how catalysts work, since they only affect the pathway, not the thermodynamics of the reaction.

3. Which acid is the weakest in aqueous solution?

A. acetic acid (Ka = 1.8 × 10^-5)

B. formic acid (Ka = 1.8 × 10^-4)

C. hydrocyanic acid (Ka = 6.2 × 10^-10)

D. nitrous acid (Ka = 4.5 × 10^-4)

Strength in water is governed by the acid dissociation constant, Ka. A larger Ka means more dissociation into H+ and A−, so a stronger acid; a very small Ka means little dissociation and a weaker acid. Hydrocyanic acid has Ka = 6.2 × 10^-10, which is far smaller than the Ka values of the other acids listed. That tiny Ka indicates it barely donates protons to water, making it the weakest acid among them. The others—with Ka on the order of 10^-5 to 10^-4—dissociate much more and are correspondingly stronger acids.

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4. During electrolysis of molten sodium iodide, what is the half-reaction occurring at the anode?

A. I- -> I+ + e-

B. Na -> Na+ + e-

C. 2 I- -> I2 + 2 e-

D. I2 + 2 e- -> 2 I-

In molten salt electrolysis, the anode is the site where oxidation occurs, so the species present as anions are the ones that get oxidized. Here, iodide (I−) is oxidized to diatomic iodine (I2), releasing electrons. The balanced half-reaction is 2 I− I2 + 2 e−. This is why iodine gas is formed at the anode. The electrons travel to the cathode, where Na+ would be reduced to Na metal: Na+ + e− Na(s). The other possibilities either propose oxidation of a species that isn't present as a metal at the anode or depict the reverse (a reduction) of iodine, which would occur at the cathode, not the anode.

5. In acidic solution, the half-reactions during electrolysis of water are:

A. Cathode: 2 H2O + 2 e− H2 + 2 OH−; Anode: 2 H2O O2 + 4 H+ + 4 e−

B. Cathode: 2 H+ + 2 e− H2; Anode: H2O 1/2 O2 + 2 H+ + 2 e−

C. Cathode: 4 H+ + 4 e− 2 H2; Anode: O2 + 4 e− 2 H2O

D. Cathode: H2O + e− 1/2 H2 + OH−; Anode: H2O H+ + e−

In acidic solution, the electrode processes during water electrolysis involve reductions that use protons (H+) at the cathode and oxidations of water at the anode. At the cathode, protons are readily available and gain electrons to form hydrogen gas: 2 H+ + 2 e− H2. At the anode, water is oxidized, producing oxygen gas and releasing protons back into the solution: 2 H2O O2 + 4 H+ + 4 e−. These two half-reactions balance the transfer of four electrons, and their sum gives the overall reaction 2 H2O 2 H2 + O2, which is the expected outcome for electrolysis of water in acidic conditions. The other possibilities misrepresent the environment or the chemistry: producing OH− at the cathode would imply a basic environment, not an acidic one; reversing the anode reaction would imply reducing O2 rather than evolving it; and using water reduction at the cathode that yields OH− again points to a basic condition. Thus, the combination of reducing protons at the cathode and oxidizing water at the anode best fits an acidic solution.

6. What is the overall balanced equation for the electrolysis of water in acidic solution?

A. 2 H2O 2 H2 + O2

B. 2 H2O H2 + O2

C. H2O H2 + 1/2 O2

D. 4 H2O 2 H2 + O2

Balancing the overall equation for electrolysis of water in acidic solution comes from adding the two electrode half-reactions after balancing electrons. At the cathode, protons are reduced to hydrogen gas: 2 H+ + 2 e- H2. At the anode, water is oxidized to oxygen gas, releasing protons and electrons: 2 H2O O2 + 4 H+ + 4 e-. To match the electrons, the cathode half-reaction is doubled: 4 H+ + 4 e- 2 H2. Now add the two half-reactions and cancel the electrons and the H+ species that appear on both sides. What remains is 2 H2O 2 H2 + O2. This is the balanced overall equation for the process in acidic solution. The other options fail because they don't balance all atoms or use fractional coefficients, which aren't appropriate for the complete, balanced equation.

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7. Calculate the standard Gibbs free energy change at 298 K for the reaction FeO_4 (s) + 4 CO (g) 3 Fe (s) + 4 CO_2 (g).

A. -14.8 kJ/mol

B. -757 kJ/mol

C. 757 kJ/mol

D. -757 kJ/mol

The concept being tested is how to get the standard Gibbs free energy change from standard Gibbs energies of formation. For a reaction, $\Delta G°rxn = \Sigma vp\Delta Gf°(products) - \Sigma vr\Delta Gf°(reactants)$. Elements in their standard state contribute zero, so Fe(s) has no effect here. For this reaction, the products are 4 CO_2, each with $\Delta Gf° \approx -394.4$ kJ/mol, giving $4 \times (-394.4) = -1577.6$ kJ. The reactants are FeO_4(s) with $\Delta Gf° \approx -1014.0$ kJ/mol and 4 CO with $\Delta Gf° \approx -137.2$ kJ/mol each, giving $(-1014.0) + 4 \times (-137.2) = -1562.8$ kJ. Therefore, $\Delta G°rxn = (-1577.6) - (-1562.8) = -14.8$ kJ/mol. So the standard Gibbs free energy change at 298 K is negative, indicating the reaction is thermodynamically favorable under standard conditions.

8. In beta decay, what happens to the atomic number Z?

A. Increases by 1

B. Increases by 2

C. Decreases by 1

D. Remains the same

In beta minus decay, a neutron in the nucleus converts into a proton, emitting an electron (the beta particle) and an antineutrino. This increases the number of protons by one, so the atomic number Z goes up by 1. The total number of nucleons stays the same, so the mass number A remains unchanged. The emitted electron carries away negative charge, but the net effect is that the nucleus gains a proton, shifting to the next higher element. (In beta plus decay, Z would decrease by 1 instead.)

9. Standard reduction potentials can be used to judge spontaneity of a redox reaction by:

A. If $E°cell < 0$ for the cell as written, the reaction is spontaneous

B. If $E°cell = 0$ for the cell as written, the reaction is spontaneous

C. If $E°cell > 0$ for the cell as written, the reaction is spontaneous

D. If E°cell is undefined, the reaction is spontaneous

The key idea is that the spontaneity of a redox reaction in standard conditions is tied to the sign of the overall standard cell potential, and to the Gibbs free energy change via $\Delta G° = -nF\ E°cell$. When E°cell is positive, $\Delta G°$ is negative, so the reaction proceeds on its own (is spontaneous) in the direction written. If $E°cell = 0$, the system is at equilibrium with no net driving force. If E°cell is negative, the reaction as written is not spontaneous; the reverse direction would be spontaneous. Therefore, the statement that the process is spontaneous when the standard cell potential for the cell as written is greater than zero is the correct criterion.

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10. Given the standard reduction potentials at 25 degrees C, what is the standard cell potential of a Ni (s) | Ni2+ (aq) || Ag+ (aq) | Ag (s) galvanic cell?

A. 0.56 V

B. 1.02 V

C. 1.35 V

D. 1.81 V

In a galvanic cell, the standard cell potential is found by taking the difference between the cathode's reduction potential and the anode's reduction potential: E°cell = E°cathode − E°anode. The Ag+/Ag couple has a higher tendency to be reduced (E° ≈ +0.80 V) than Ni2+/Ni (E° ≈ −0.23 V), so silver is reduced at the cathode and nickel is oxidized at the anode. Balancing electrons gives Ni(s) Ni2+(aq) + 2 e− and 2 Ag+(aq) + 2 e− 2 Ag(s). Therefore E°cell = +0.80 − (−0.23) ≈ 1.03 V, which rounds to about 1.02 V. So the standard cell potential is approximately 1.02 V.

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Next Steps

Congratulations on reaching the final section of this guide. You've taken a meaningful step toward passing your certification exam and advancing your career.

As you continue preparing, remember that consistent practice, review, and self-reflection are key to success. Make time to revisit difficult topics, simulate exam conditions, and track your progress along the way.

If you need help, have suggestions, or want to share feedback, we'd love to hear from you. Reach out to our team at hello@examzify.com.

Or visit your dedicated course page for more study tools and resources:

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We wish you the very best on your exam journey. You've got this!

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