ANSWER KEY

General Chemistry 2: Practice Exam

Quick answers first, then a worked solution for every problem.

#	Topic	Answer
1	Freezing point depression	−2.55 °C
2	Rate law from initial rates	Rate = $k[A]^2$, $k = 0.040\ M^{-1}s^{-1}$
3	First-order kinetics and half-life	(a) 1.1×10^3 s (about 19 min) (b) 2.6×10^3 s (about 43 min)
4	Activation energy (Arrhenius)	53.6 kJ/mol
5	Reaction quotient Q vs K	Q = 2.5 > K, shifts left
6	ICE table	[HI] = 0.311 M, $[H_2] = [I_2] = 0.044$ M
7	Le Chatelier's principle	(a) left (b) right (c) right (d) no shift
8	pH of a strong base	pH = 12.00
9	pH of a weak acid	pH = 2.87, 1.3% ionized
10	pH of a salt solution	pH = 8.23
11	Buffers	(a) 4.95 (b) 4.89
12	Titration of a weak acid	(a) 4.74 (b) 8.72
13	K_{sp} and the common-ion effect	(a) 2.1×10^{-4} M (b) 3.9×10^{-7} M
14	Gibbs free energy	(a) ΔG° = +130.4 kJ, no (b) 1111 K (838 °C) (c) $K \approx 1.4 \times 10^{-23}$
15	Electrochemistry	(a) 1.10 V (b) −212 kJ (c) 1.04 V

1. Freezing point depression

10.0 g of NaCl dissolves in 250. g of water. What is the freezing point? (K_f of water = 1.86 °C/m.)

1. Moles NaCl = 10.0 / 58.44 = 0.1711 mol. Molality = 0.1711 / 0.250 kg = 0.6845 m.
2. NaCl splits into 2 ions, so i = 2.
3. $\Delta T_f = i \times K_f \times m = 2 \times 1.86 \times 0.6845 = 2.55$ °C. Freezing point = **−2.55 °C**.

> **How to explain it:** Colligative properties count particles, not identity. Forgetting i = 2 cuts the answer in half, and molality uses kilograms of solvent, not liters of solution.

2. Rate law from initial rates

For A + B → products:

Experiment	[A] (M)	[B] (M)	Initial rate (M/s)
1	0.10	0.10	4.0×10^{-4}
2	0.20	0.10	1.6×10^{-3}
3	0.10	0.30	4.0×10^{-4}

Find the rate law and k.

1. Experiments 1 and 2: [A] doubles and the rate goes up 4 times, so A is **second order**.
2. Experiments 1 and 3: [B] triples and the rate doesn't change, so B is **zero order**.
3. Rate = $k[A]^2$. Using experiment 1: $k = 4.0 \times 10^{-4} / (0.10)^2$ = **$0.040\ M^{-1}s^{-1}$**.

> **How to explain it:** Pick two experiments where only one concentration changes. Orders come from data, never from the coefficients in the equation.

3. First-order kinetics and half-life

N_2O_5 decomposes by first-order kinetics with $k = 6.2 \times 10^{-4}\ s^{-1}$. (a) What is the half-life? (b) How long does it take for $[N_2O_5]$ to drop from 0.50 M to 0.10 M?

1. $t_{1/2} = 0.693 / k = 0.693 / 6.2 \times 10^{-4} =$ **1.1×10^3 s** (about 19 minutes).
2. $\ln([A]_0/[A]) = kt$, so $t = \ln(0.50/0.10) / 6.2 \times 10^{-4} = 1.609 / 6.2 \times 10^{-4} =$ **2.6×10^3 s (about 43 minutes)**.

> **How to explain it:** A first-order half-life doesn't depend on concentration. Students often put the ratio upside down inside the ln and get a negative time.

4. Activation energy (Arrhenius)

A reaction's rate doubles when the temperature rises from 300 K to 310 K. Find E_a.

1. $\ln(k_2/k_1) = (E_a/R)(1/T_1 - 1/T_2)$.
2. $\ln 2 = (E_a / 8.314)(1/300 - 1/310) = (E_a / 8.314)(1.075 \times 10^{-4})$.
3. $E_a = 0.6931 \times 8.314 / 1.075 \times 10^{-4} = 5.36 \times 10^4$ J/mol = **53.6 kJ/mol**.

> **How to explain it:** This is the "rate doubles every 10 degrees" rule of thumb, worked out. Use R = 8.314 J/(mol·K) and convert to kJ at the end.

5. Reaction quotient Q vs K

$N_2 + 3\ H_2 \rightleftharpoons 2\ NH_3$, $K_c = 0.105$. A mixture has $[N_2] = 0.50$ M, $[H_2] = 0.20$ M and $[NH_3] = 0.10$ M. Which way does it shift?

1. $Q = [NH_3]^2 / ([N_2][H_2]^3) = (0.10)^2 / ((0.50)(0.20)^3) = 0.010 / 0.0040 = 2.5$.
2. Q (2.5) > K (0.105): too much product, so the reaction **shifts left**, toward N_2 and H_2.

> **How to explain it:** Q is a snapshot and K is the goal. Q bigger than K means there's too much on top, so the reaction runs backward. Students forget to cube $[H_2]$.

6. ICE table

$H_2 + I_2 \rightleftharpoons 2\ HI$, $K_c = 49.0$. You start with 0.200 M H_2 and 0.200 M I_2. Find the equilibrium concentrations.

1. ICE: H_2 and I_2 each go 0.200 → 0.200 − x; HI goes 0 → 2x.
2. $49.0 = (2x)^2 / (0.200 - x)^2$. Take the square root: $7.00 = 2x / (0.200 - x)$.
3. $1.400 - 7.00x = 2x$, so $x = 0.1556$.
4. **$[HI] = 0.311$ M, $[H_2] = [I_2] = 0.044$ M.** Check: $(0.311)^2 / (0.044)^2 \approx 49$.

> **How to explain it:** Look for a perfect square before reaching for the quadratic formula. The change for HI is 2x because of the coefficient, and that's the step students miss.

7. Le Chatelier's principle

$N_2 + 3\ H_2 \rightleftharpoons 2\ NH_3$, $\Delta H = -92$ kJ. Which way does it shift if you (a) raise the temperature, (b) add H_2, (c) halve the volume, (d) add a catalyst?

1. (a) **Left.** The reaction is exothermic, so heat acts like a product. Raising T also makes K smaller.
2. (b) **Right**, to use up the added H_2.
3. (c) **Right**, toward fewer moles of gas (4 on the left, 2 on the right).
4. (d) **No shift.** A catalyst speeds up both directions equally.

> **How to explain it:** Treat heat as a reactant or product, then push the opposite way. Only temperature changes K; the other changes just move the mixture back to the same K.

8. pH of a strong base

What is the pH of 0.0050 M $Ba(OH)_2$?

1. Each $Ba(OH)_2$ gives 2 OH^-: $[OH^-] = 0.010$ M.
2. $pOH = -\log(0.010) = 2.00$; $pH = 14.00 - 2.00 =$ **12.00**.

> **How to explain it:** Count the hydroxides. Using 0.0050 M directly gives 11.70, a common wrong answer on multiple choice.

9. pH of a weak acid

Find the pH and percent ionization of 0.10 M acetic acid ($K_a = 1.8 \times 10^{-5}$).

1. $K_a = x^2 / (0.10 - x) \approx x^2 / 0.10$, so $x = \sqrt{(1.8 \times 10^{-6})} = 1.34 \times 10^{-3}$ M.
2. pH = $-\log(1.34 \times 10^{-3})$ = **2.87**.
3. Percent ionization = $1.34 \times 10^{-3} / 0.10 \times 100$ = **1.3%**. That's under 5%, so dropping x was fine.

> **How to explain it:** Weak means x is small, so check the 5% rule after. Treating it like a strong acid gives pH 1.00.

10. pH of a salt solution

What is the pH of 0.20 M NaF? (K_a of HF = 6.8×10^{-4}.)

1. F^- is a weak base: $F^- + H_2O \rightleftharpoons HF + OH^-$.
2. $K_b = K_w / K_a = 1.0 \times 10^{-14} / 6.8 \times 10^{-4} = 1.5 \times 10^{-11}$.
3. $[OH^-] = \sqrt{(1.5 \times 10^{-11} \times 0.20)} = 1.7 \times 10^{-6}$ M, so pOH = 5.77 and **pH = 8.23**.

> **How to explain it:** Ask which ion came from a weak acid or base. Na^+ does nothing, F^- grabs H^+ from water, so the solution is slightly basic.

11. Buffers

A 1.00 L buffer has 0.25 M acetic acid and 0.40 M sodium acetate ($pK_a = 4.74$). (a) What is its pH? (b) What is the pH after adding 0.020 mol HCl?

1. (a) pH = $pK_a + \log([A^-]/[HA])$ = $4.74 + \log(0.40/0.25)$ = **4.95**.
2. (b) HCl turns acetate into acetic acid: acetate 0.40 → 0.38 mol, acid 0.25 → 0.27 mol.
3. pH = $4.74 + \log(0.38/0.27)$ = **4.89**. It drops only 0.06.

> **How to explain it:** Do the reaction first (stoichiometry), then the Henderson–Hasselbalch equation. Putting base over acid upside down is the mistake to watch for.

12. Titration of a weak acid

25.0 mL of 0.100 M acetic acid is titrated with 0.100 M NaOH. Find the pH (a) at 12.5 mL NaOH and (b) at 25.0 mL NaOH. (K_a = 1.8 × 10^{-5}.)

1. (a) 12.5 mL is halfway to equivalence, so [HA] = [A^-] and **pH = pK_a = 4.74**.
2. (b) At equivalence all acid is now acetate: 0.00250 mol in 50.0 mL = 0.0500 M.
3. K_b = 1.0 × 10^{-14} / 1.8 × 10^{-5} = 5.6 × 10^{-10}; [OH^-] = √(5.6 × 10^{-10} × 0.0500) = 5.3 × 10^{-6} M.
4. pOH = 5.28, so **pH = 8.72**.

> **How to explain it:** Half-equivalence pH equals pK_a, the most tested fact on the curve. A weak acid's equivalence point is above 7, not exactly 7, and the volume doubles to 50.0 mL.

13. K_{sp} and the common-ion effect

K_{sp} of CaF_2 = 3.9 × 10^{-11}. Find its molar solubility (a) in pure water and (b) in 0.010 M NaF.

1. (a) $CaF_2 \rightleftharpoons Ca^{2+} + 2\ F^-$, so $K_{sp} = (s)(2s)^2 = 4s^3$. $s = (3.9 \times 10^{-11} / 4)^{1/3}$ = **2.1 × 10^{-4} M**.
2. (b) [F^-] ≈ 0.010 M from NaF, so $K_{sp} = s(0.010)^2$ and s = 3.9 × 10^{-11} / 1.0 × 10^{-4} = **3.9 × 10^{-7} M**, about 550 times less.

> **How to explain it:** Write the ICE line with 2s for fluoride, then square it. The common ion is already there, so it pushes the equilibrium left and less dissolves.

14. Gibbs free energy

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$, ΔH° = +178.3 kJ, ΔS° = +160.5 J/K. (a) Is it spontaneous at 25 °C? (b) Above what temperature is it spontaneous? (c) What is K at 25 °C?

1. (a) ΔG° = ΔH° − TΔS° = 178.3 − (298.15)(0.1605) = **+130.4 kJ**, so not spontaneous.
2. (b) ΔG° = 0 when T = ΔH°/ΔS° = 178.3 / 0.1605 = **1111 K (838 °C)**. It's spontaneous above that.
3. (c) ln K = −ΔG°/RT = −130,400 / (8.314 × 298.15) = −52.6, so **K ≈ 1.4 × 10^{-23}**.

> **How to explain it:** Convert ΔS to kJ/K (or ΔH to J) before you subtract. Mixing units is how students get ΔG of about −47,000.

15. Electrochemistry

A galvanic cell uses Zn | Zn^{2+} and Cu^{2+} | Cu. $E°(Cu^{2+}/Cu)$ = +0.34 V, $E°(Zn^{2+}/Zn)$ = −0.76 V. Find (a) $E°_{cell}$, (b) $\Delta G°$, and (c) E_{cell} when $[Zn^{2+}]$ = 1.0 M and $[Cu^{2+}]$ = 0.010 M.

1. (a) Cu^{2+} is reduced (cathode), Zn is oxidized (anode). $E°_{cell}$ = 0.34 − (−0.76) = **1.10 V**.
2. (b) $\Delta G° = -nFE°$ = −(2)(96,485)(1.10) = **−212 kJ**.
3. (c) Nernst: $E = E° - (0.0592/n) \log Q$ = 1.10 − (0.0592/2) log(1.0/0.010) = **1.04 V**.

> **How to explain it:** $E°_{cell}$ = cathode minus anode, with both written as reductions. Don't multiply E° by coefficients. Less Cu^{2+} means less push to reduce, so the voltage drops.