ExamFitter

examfitter.com · Free AP Chemistry quizzes, revision notes and worksheets

Topic 9.5: Free Energy and Equilibrium

Worksheet · AP Chemistry

UNIT 9 · THERMODYNAMICS AND ELECTROCHEMISTRY

On ExamFitter.com: Topic 9.5 Revision Notes · Topic 9.5 Quiz · Previous Worksheet: Topic 9.4 · Next Worksheet: Topic 9.6 · Unit 9: All Topics · AP Chemistry Home

How to Use This Worksheet

Work in order. Section A checks the ideas: reading the sign of ΔG° from K, separating ΔG from ΔG°, the units inside the equation, and the 5.7 kJ estimation rule. Section B practices the calculations themselves: K from ΔG°, ΔG° from K, the effect of temperature on the conversion, direction from Q, and what doubling ΔG° really does to K. Section C is exam style: a temperature study of ammonia synthesis, a full reaction quotient problem, and two reactions compared through their ΔG° values.

Reference facts. R = 8.314 J mol^{-1} K^{-1} throughout, and RT = 2.478 kJ mol^{-1} at 298 K. Temperatures are always in kelvin. ΔG° values are quoted in kJ mol^{-1} and must be converted to joules before they enter $\Delta G^\circ = -RT \ln K$. One system recurs across the sheet: $N_2O_4(g) \rightleftharpoons 2\,NO_2(g)$, which has K = 0.15 at 298 K and ΔG° = +4.7 kJ mol^{-1}, with Q built from partial pressures in atm as $Q = P(NO_2)^2 / P(N_2O_4)$.

One working habit ties the sheet together: sign first, units second, arithmetic third. Compare K with 1, or Q with K, and name the verdict before calculating anything. Then convert: joules and kelvin. Then calculate. Finally, check: every K calculated from a ΔG° is converted back to confirm the starting value, and every direction claim is confirmed twice, once by the Q against K comparison and once by the sign of ΔG.

The answer key gives the full working for every part. Do not look at the answer key until you have attempted every question.

© 2026 ExamFitter.com · Free AP Chemistry worksheets

Section A: Recall and Concepts

A1. (a) For each value of K at 298 K, state the sign of ΔG° and which side, reactants or products, is favored at equilibrium: $K = 4.2 \times 10^{8}$; K = 0.020; K = 1. (b) A reaction has $\Delta G^\circ = +35$ kJ mol^{-1} at 298 K, giving $K = 7.3 \times 10^{-7}$. Explain why the equilibrium mixture still contains some product, even though K is so small.

A2. (a) Write the equation connecting ΔG° and K, defining every symbol and giving the unit of each quantity as it is used in the equation. (b) State the two conversions that must be applied to typical data, a ΔG° in kJ mol^{-1} and a temperature in °C, before the equation is used. (c) The equation uses the natural logarithm, ln. State the base 10 form of the equation and the extra factor it requires.

A3. (a) Distinguish ΔG° from ΔG: what each quantity describes, and what fixes the value of each. (b) A reaction has $\Delta G^\circ = -40$ kJ mol^{-1}. A student claims the reaction must proceed forward from any starting mixture. Judge the claim. (c) The same reaction is placed in a vessel containing almost pure product. State the sign of ΔG for that mixture and the direction of the net reaction.

A4. (a) State the 5.7 kJ rule at 298 K precisely, including what stays fixed as ΔG° changes by that step. (b) Use the rule to estimate K at 298 K for $\Delta G^\circ = -17.1$ kJ mol^{-1} and for $\Delta G^\circ = +5.7$ kJ mol^{-1}. (c) Use the rule to estimate ΔG° at 298 K for a reaction with $K = 1.0 \times 10^{5}$. (d) At 500 K the step size RT ln 10 is 9.6 kJ mol^{-1}, not 5.7. Explain in one sentence why the rule cannot be carried to other temperatures unchanged.

© 2026 ExamFitter.com · Free AP Chemistry worksheets

A5. Judge each claim, giving a reason. (a) "ΔG° = 0 means the reaction has stopped." (b) "A catalyst raises K, because it helps the reaction reach equilibrium." (c) "Changing the initial concentrations of a mixture changes ΔG°." (d) "Heating an exothermic reaction increases its equilibrium constant."

Section B: Core Practice

B1. At 298 K, a reaction has K = 0.020. (a) Calculate ΔG° in kJ mol^{-1}, showing each step. (b) State which side is favored at equilibrium. (c) Check your answer by converting your ΔG° back into K.

B2. A reaction has ΔG° = +15 kJ mol^{-1} at 298 K. (a) Calculate K. (b) Check your answer by converting your K back into ΔG°. (c) State which side is favored at equilibrium, and whether any product forms at all.

B3. A reaction has ΔG° = −11.4 kJ mol^{-1} at 298 K. (a) Estimate K using the 5.7 kJ rule, with no logarithm calculation. (b) Calculate K using ΔG° = −RT ln K. (c) State whether the estimate and the calculation agree, and why the estimate works at all.

B4. At 1000 K, a reaction has ΔG° = −8.314 kJ mol^{-1}. (a) Calculate K. (b) A second reaction at 1000 K has ΔG° = −16.628 kJ mol^{-1}. State ln K for it, and hence K, explaining the shortcut. (c) The first reaction's ΔG°, −8.314 kJ mol^{-1}, is moved to 298 K unchanged. Calculate the K it gives there, and explain why the same energy produces a larger K at the lower temperature.

© 2026 ExamFitter.com · Free AP Chemistry worksheets

B5. For $N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$ at 298 K, K = 0.15 and $\Delta G° = +4.7$ kJ mol^{-1}. (a) A mixture has $P(N_2O_4)$ = 0.50 atm and $P(NO_2)$ = 0.10 atm. Calculate Q, compare it with K, and state the direction of the net reaction. (b) Calculate ΔG for the mixture in (a), and confirm that its sign agrees with your direction. (c) A second mixture has $P(N_2O_4)$ = 0.10 atm and $P(NO_2)$ = 0.50 atm. Calculate Q and ΔG, and state the direction.

B6. At 500 K, reaction 1 has K = 50. (a) Calculate ΔG° for reaction 1. (b) Reaction 2, also at 500 K, has a ΔG° twice as negative as reaction 1. Calculate K for reaction 2, showing the ln step explicitly. (c) A student answers 100 for part (b), reasoning that "everything doubled". Explain the student's error.

Section C: Exam Style

C1. Ammonia synthesis, $N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g)$, has $\Delta H° = -92$ kJ mol^{-1} and $\Delta S° = -199$ J mol^{-1} K^{-1}, both treated as constant over the range studied. (a) Calculate ΔG° at 298 K, and hence K at 298 K. (b) Calculate ΔG° at 700 K, and hence K at 700 K. (c) State the temperature at which K = 1, and verify it by substituting that temperature into $\Delta G° = \Delta H° - T\Delta S°$. (d) Explain why heating this reaction lowers K, naming the two equations whose verdicts agree.

© 2026 ExamFitter.com · Free AP Chemistry worksheets

C2. For $N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$ at 298 K, K = 0.15 and ΔG° = +4.7 kJ mol^{-1}. A mixture is prepared with $P(N_2O_4)$ = 0.80 atm and $P(NO_2)$ = 0.20 atm. (a) Calculate Q for the mixture. (b) State the direction of the net reaction, using only the comparison of Q with K. (c) Calculate ΔG for the mixture, and confirm that its sign agrees with (b). (d) State the values of Q and ΔG once the mixture reaches equilibrium, and the value of ΔG° at that moment.

C3. Reactions X and Y both run at 298 K. Reaction X has $\Delta G^\circ = -12$ kJ mol^{-1} and reaction Y has $\Delta G^\circ = -25$ kJ mol^{-1}. (a) Calculate K for each reaction. (b) Calculate the ratio K(Y) / K(X), and check it against the 5.7 kJ rule applied to the 13 kJ mol^{-1} difference between the two ΔG° values. (c) A student argues: "Y's ΔG° is about twice X's, so Y's equilibrium constant is about twice X's." Judge the argument using your answers.

© 2026 ExamFitter.com · Free AP Chemistry worksheets

Answer Key

A1. (a) $K = 4.2 \times 10^{8}$: ln K is positive and large, so $\Delta G°$ is negative (-49.2 kJ mol^{-1}); products are favored. $K = 0.020$: ln K is negative, so $\Delta G°$ is positive ($+9.7$ kJ mol^{-1}); reactants are favored. $K = 1$: ln 1 = 0, so $\Delta G° = 0$; neither side is favored. (b) K is a ratio, products over reactants at equilibrium. A value of 7.3×10^{-7} means the product side is present at about seven ten-millionths of the reactant side's measure, not at zero. Reactants are overwhelmingly favored, but the equilibrium is a balance, not an absence: some product is always present.

A2. (a) $\Delta G° = -RT \ln K$, where $\Delta G°$ is the standard free energy change in J mol^{-1} while inside this equation, $R = 8.314$ J mol^{-1} K^{-1}, T is the absolute temperature in kelvin, and K is the equilibrium constant for the reaction at that temperature. (b) $\Delta G°$ must be converted from kJ mol^{-1} to J mol^{-1} by multiplying by 1000, because R carries joules. The temperature must be converted from °C to kelvin by adding 273. (c) $\Delta G° = -2.303\ RT \log K$. The base 10 form needs the extra factor 2.303, the value of ln 10, because log and ln differ by exactly that factor.

A3. (a) $\Delta G°$ is the free energy change when every species is in its standard state. For a given reaction it depends only on temperature, and it states which side is favored at equilibrium. ΔG is the free energy change under the actual conditions in a real mixture; it changes as the composition changes, and its sign states the direction the mixture moves toward equilibrium. (b) The claim is false. $\Delta G°$ describes standard conditions and the equilibrium position only. Direction is set by the sign of ΔG for the actual mixture, and a mixture loaded with product has a positive ΔG and runs in reverse, whatever $\Delta G°$ says. (c) ΔG is positive, and the net reaction proceeds in reverse, converting product back into reactants. As it does, Q falls toward K, and the motion stops when $Q = K$ and ΔG reaches zero.

A4. (a) At 298 K, $RT \ln 10 = 5.7$ kJ mol^{-1}, so every change of 5.7 kJ mol^{-1} in $\Delta G°$ multiplies or divides K by a factor of exactly 10, with more negative $\Delta G°$ giving the larger K. (b) -17.1 kJ mol^{-1} is three steps of -5.7, so $K \approx 10^{3} = 1.0 \times 10^{3}$ (exact value 994). $+5.7$ kJ mol^{-1} is one step in the positive direction, so $K \approx 10^{-1} = 0.10$ (exact value 0.100). (c) $K = 1.0 \times 10^{5}$ is five powers of ten above 1, so $\Delta G° \approx -5 \times 5.7 = -28.5$ kJ mol^{-1} (exact value -28.5 kJ mol^{-1}). (d) The step size is $RT \ln 10$, and T sits inside it: raise the temperature and the same factor of 10 in K costs more energy, 9.6 kJ mol^{-1} at 500 K instead of 5.7.

A5. (a) False. $\Delta G° = 0$ means $K = 1$: the system is at equilibrium with forward and reverse reactions running at equal rates. Nothing has stopped; the composition has stopped changing. (b) False. A catalyst lowers the activation energy of both directions equally and changes only how quickly equilibrium is reached. K and $\Delta G°$ are properties of the reactants and products, so both are untouched. (c) False. $\Delta G°$ depends only on temperature for a given reaction. Changing initial concentrations changes Q, and therefore ΔG, which is exactly why mixtures starting from different compositions can move in different directions under one fixed $\Delta G°$. (d) False. For an exothermic

© 2026 ExamFitter.com · Free AP Chemistry worksheets

reaction, raising the temperature shifts the equilibrium toward reactants, so K falls; ammonia synthesis drops from $K = 5.4 \times 10^{5}$ at 298 K to 3.0×10^{-4} at 700 K.

B1. (a) $\ln(0.020) = -3.91$. $\Delta G° = -RT \ln K = -(8.314\ \text{J mol}^{-1}\ \text{K}^{-1})(298\ \text{K})(-3.91) = +9690\ \text{J mol}^{-1} = +9.7\ \text{kJ mol}^{-1}$. (b) Reactants are favored: $\Delta G°$ is positive, matching $K < 1$. (c) $\ln K = -\Delta G° / RT = -9690 / 2477.6 = -3.91$, so $K = e^{-3.91} = 0.020$. The starting value is recovered, so the calculation is consistent.

B2. (a) $\ln K = -\Delta G° / RT = -15{,}000 / 2477.6 = -6.05$, so $K = e^{-6.05} = 2.4 \times 10^{-3}$. (b) $\Delta G° = -RT \ln K = -(2477.6)(\ln 2.4 \times 10^{-3}) = -(2477.6)(-6.05) = +15{,}000\ \text{J mol}^{-1} = +15\ \text{kJ mol}^{-1}$, the starting value. (c) Reactants are favored, since K is far below 1, but product does form: at equilibrium the products sit at about 2.4 thousandths of the reactants' measure, small but not zero.

B3. (a) $11.4 / 5.7 = 2$ steps in the negative direction, so $K \approx 10^{2} = 1.0 \times 10^{2}$. (b) $\ln K = -\Delta G° / RT = 11{,}400 / 2477.6 = 4.60$, so $K = e^{4.60} = 99.6 \approx 1.0 \times 10^{2}$. (c) They agree, to the precision the rule claims. The rule works because the equation is logarithmic: equal steps in $\Delta G°$ are equal ratios in K, and at 298 K the step that produces a ratio of 10 is exactly $RT \ln 10 = 5.7\ \text{kJ mol}^{-1}$.

B4. (a) $\ln K = -\Delta G° / RT = 8314 / (8.314 \times 1000) = 1.00$, so $K = e^{1} = 2.7$. (b) Doubling $\Delta G°$ doubles $\ln K$: $\ln K = 2.00$, so $K = e^{2} = 7.4$. The shortcut works because $\Delta G°$ enters the equation only through the ratio $-\Delta G° / RT$, so scaling $\Delta G°$ scales $\ln K$ by the same factor. (c) $\ln K = 8314 / (8.314 \times 298) = 8314 / 2477.6 = 3.36$, so $K = e^{3.36} = 28.7 \approx 29$. The energy is compared against RT: at the lower temperature RT is smaller, so the same $-8.314\ \text{kJ mol}^{-1}$ is a larger multiple of RT, giving a larger $\ln K$ and so a larger K.

B5. (a) $Q = P(NO_2)^2 / P(N_2O_4) = (0.10)^2 / 0.50 = 0.020$. $Q < K$ ($0.020 < 0.15$): the mixture is short of product, so the net reaction proceeds forward. (b) $\Delta G = \Delta G° + RT \ln Q = +4.70 + (2.478)(\ln 0.020) = +4.70 + (2.478)(-3.91) = +4.70 - 9.69 = -5.0\ \text{kJ mol}^{-1}$. Negative, which means forward, agreeing with (a). (c) $Q = (0.50)^2 / 0.10 = 2.5$. $\Delta G = +4.70 + (2.478)(\ln 2.5) = +4.70 + (2.478)(0.916) = +4.70 + 2.27 = +7.0\ \text{kJ mol}^{-1}$. $Q > K$ and ΔG is positive: both tests say the net reaction proceeds in reverse.

B6. (a) $\Delta G°_1 = -RT \ln K = -(8.314)(500)(\ln 50) = -(4157)(3.912) = -16{,}262\ \text{J mol}^{-1} = -16.3\ \text{kJ mol}^{-1}$. (b) $\Delta G°_2 = 2 \times (-16{,}262) = -32{,}524\ \text{J mol}^{-1}$. $\ln K_2 = -\Delta G°_2 / RT = 32{,}524 / 4157 = 7.82$, which is exactly $2 \times \ln 50$. So $K_2 = e^{7.82} = 2500$, equivalently $K_2 = 50^{2}$. (c) K is an exponential function of $\Delta G°$, not a linear one. Doubling $\Delta G°$ doubles the exponent $\ln K$, and doubling an exponent squares the number: 50 becomes $50^{2} = 2500$, not 100.

C1. (a) Convert $\Delta S°$: $-199\ \text{J mol}^{-1}\ \text{K}^{-1} = -0.199\ \text{kJ mol}^{-1}\ \text{K}^{-1}$. $\Delta G° = \Delta H° - T\Delta S° = -92 - (298)(-0.199) = -92 + 59.3 = -32.7\ \text{kJ mol}^{-1}$. Then $\ln K = 32{,}700 / 2477.6 = 13.2$, so $K = e^{13.2} = 5.4 \times 10^{5}$. Products are strongly favored at 298 K. (b) $\Delta G° = -92 - (700)(-0.199) = -92 + 139.3 = +47.3\ \text{kJ mol}^{-1}$. RT at 700 K is $8.314 \times 700 = 5820\ \text{J mol}^{-1}$, so $\ln K = -47{,}300 / 5820 = -8.13$, and $K = e^{-8.13} = 3.0 \times 10^{-4}$. The equilibrium has swung to the reactants. (c) $K = 1$ requires $\Delta G° = 0$, so $T = \Delta H° / \Delta S° = 92{,}000 / 199 = 462$ K. Verification: $\Delta G° = -92 - (462)(-0.199) = -92 + 91.9 \approx 0\ \text{kJ mol}^{-1}$, zero to the precision of the data. (d) In $\Delta G° = \Delta H° - T\Delta S°$, this reaction's $\Delta S°$ is negative, so the $-T\Delta S°$ term is positive and

© 2026 ExamFitter.com · Free AP Chemistry worksheets

grows with T, dragging ΔG° upward through zero. ΔG° = −RT ln K then forces ln K, and so K, down as ΔG° rises. Le Chatelier's principle agrees from the other side: the reaction is exothermic, heat is effectively a product, and adding heat shifts the equilibrium toward reactants.

C2. (a) $Q = P(NO_2)^2 / P(N_2O_4) = (0.20)^2 / 0.80 = 0.040 / 0.80 = 0.050$. (b) $Q = 0.050 < K = 0.15$: the mixture holds too little NO_2 relative to equilibrium, so the net reaction proceeds forward. (c) $\Delta G = \Delta G° + RT \ln Q = +4.70 + (2.478)(\ln 0.050) = +4.70 + (2.478)(-3.00) = +4.70 - 7.42 = -2.7$ kJ mol^{-1}. Negative, so forward, agreeing with (b). (d) At equilibrium Q has risen to equal K: $Q = 0.15$, and $\Delta G = 0$. ΔG° is unchanged at +4.7 kJ mol^{-1}, because the temperature has not changed; ΔG° never moves during an approach to equilibrium.

C3. (a) Reaction X: $\ln K = 12{,}000 / 2477.6 = 4.84$, so $K(X) = e^{4.84} = 127 \approx 1.3 \times 10^2$. Reaction Y: $\ln K = 25{,}000 / 2477.6 = 10.09$, so $K(Y) = e^{10.09} = 2.4 \times 10^4$. (b) $K(Y) / K(X) = 2.4 \times 10^4 / 127 = 190$. By the rule: the difference is 25 − 12 = 13 kJ mol^{-1}, and 13 / 5.7 = 2.28 powers of ten, so the ratio should be $10^{2.28} \approx 190$. The two routes agree. (c) The argument fails. Y's ΔG° is about twice X's, but Y's K is about 190 times X's, not twice. K responds exponentially to ΔG°: equal additions to ΔG° multiply K by a fixed factor, here about $10^{2.28}$ for the 13 kJ mol^{-1} gap, so a doubling of ΔG° produces a far more than doubled K.

Continue Revising

Topic 9.5 Revision Notes

Topic 9.5 Quiz: Free Energy and Equilibrium

Previous Worksheet: Topic 9.4 Thermodynamic and Kinetic Control

Next Worksheet: Topic 9.6 Free Energy of Dissolution

© 2026 ExamFitter.com · Free AP Chemistry worksheets