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Topic 7.3: Reaction Quotient and Equilibrium Constant

Worksheet · AP Chemistry

UNIT 7 · EQUILIBRIUM

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How to Use This Worksheet

Work in order. Section A checks the building blocks: writing Kc expressions, leaving pure solids and liquids out, the difference between Q and K, the three direction cases, and what happens to Q when every concentration is scaled. Section B works one system at a time: three mixtures of hydrogen, iodine and hydrogen iodide against Kc = 64, the ways of rewriting that constant, a limestone decomposition, and a Kp calculation. Section C is exam style: a water gas shift mixture with a sudden volume change, a target constant built from two given reactions, and five claims to judge.

Reference facts. Kc = products over reactants, each concentration raised to its coefficient; pure solids and liquids are omitted. Q uses current concentrations, K uses equilibrium concentrations; Q < K means net forward, Q > K net reverse, Q = K at equilibrium. Kp = Kc(RT)^Δn, with Δn = gas moles of products minus reactants. Useful values: the HI system has Kc = 64; the SO_2 system has Kc = 2.5×10^2 at 1000 K and Kp = 3.0; the limestone system has Kc = 0.050.

Three habits protect every answer. Write the expression first, then the substitution line, then the result. Put coefficients in as powers, never as multipliers. And compare Q with K before stating any direction.

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In every calculation, write the expression first, then the substitution line with the numbers in, then the result: the substitution is where the powers and the omitted species are checked. Concentrations are in mol L^{-1} and partial pressures in atm throughout.

For the Kp work, use R = 0.08206 L atm mol^{-1} K^{-1} with temperatures in kelvin, and count only gas species in Δn: solids and liquids contribute no gas moles. Equilibrium constants are quoted as pure numbers, without units, in every question.

Do not look at the answer key until you have attempted every question.

Section A: Recall and Concepts

A1. (a) Write the Kc expression for $2NO(g) + Cl_2(g) \rightleftharpoons 2NOCl(g)$. (b) Write the Kc expression for $N_2O_4(g) \rightleftharpoons 2NO_2(g)$. (c) State the two rules that decide the form of each expression.

A2. (a) Explain why $CaCO_3(s)$ and $CaO(s)$ do not appear in the Kc expression for $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$. (b) Write Kc for that reaction. (c) Write Kc for $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$, and name the species omitted.

A3. (a) State the difference between Q and K for the same reaction at the same temperature. (b) Describe what happens to the value of Q as a mixture approaches equilibrium, covering both possible starting sides. (c) State the one change that alters the value of K for a given reaction.

A4. For each case, state the net direction and what happens to the product concentrations as the system moves: (a) $Q < K$; (b) $Q > K$; (c) $Q = K$.

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A5. (a) For $2NO(g) + Cl_2(g) \rightleftharpoons 2NOCl(g)$, every concentration in a mixture is doubled. Work out the factor by which Q changes, showing the power totals in the numerator and the denominator. (b) Repeat the calculation for $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$. (c) State the feature of a balanced equation that decides whether doubling every concentration changes Q at all.

Section B: Core Practice

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ has Kc = 64 at the temperature of the mixtures in B1 to B4. The three mixtures are separate vessels held at that same temperature, not one vessel watched over time, so each mixture is judged on its own value of Q.

B1. A mixture has $[H_2] = 0.125$ mol L^{-1}, $[I_2] = 0.125$ mol L^{-1} and $[HI] = 1.00$ mol L^{-1}. (a) Write the expression for Q. (b) Substitute the concentrations and calculate Q. (c) Compare Q with Kc and state what the result shows about this mixture.

B2. A second mixture has $[H_2] = 0.20$ mol L^{-1}, $[I_2] = 0.20$ mol L^{-1} and $[HI] = 0.80$ mol L^{-1}. (a) Calculate Q, showing the substitution. (b) State the net direction, with the comparison that decides it. (c) State which concentrations rise and which fall as the system moves.

B3. A third mixture has $[H_2] = 0.050$ mol L^{-1}, $[I_2] = 0.050$ mol L^{-1} and $[HI] = 0.60$ mol L^{-1}. (a) Calculate Q, showing the substitution. (b) State the net direction, with the comparison that decides it. (c) State the value Q will reach when the movement stops, and describe what the concentrations are doing at that point.

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B4. Use Kc = 64 for $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ as written. (a) State Kc for $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$, naming the rule used. (b) State Kc for $\frac{1}{2}H_2(g) + \frac{1}{2}I_2(g) \rightleftharpoons HI(g)$, naming the rule used. (c) State Kc for $2H_2(g) + 2I_2(g) \rightleftharpoons 4HI(g)$, naming the rule used.

B5. $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ has Kc = 0.050 at a certain temperature. Both solids are present in each vessel. (a) Write the expression for Q. (b) A vessel has $[CO_2] = 0.020$ mol L^{-1}. Calculate Q, state the net direction, and state what happens to the amount of $CaCO_3(s)$. (c) A second vessel has $[CO_2] = 0.085$ mol L^{-1}. Calculate Q, state the net direction, and state what happens to the amount of CaO(s). (d) Extra $CaCO_3(s)$ is added to the first vessel while $[CO_2]$ is still 0.020 mol L^{-1}. State the new value of Q and explain the answer. (e) State the value of $[CO_2]$ in each of the two vessels once all movement has stopped, and the condition that must hold for that value to be reached.

B6. For $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ at 1000 K, $Kc = 2.5 \times 10^2$. (a) Write the Kp expression. (b) Calculate Δn for the reaction as written. (c) Calculate Kp, using Kp = Kc(RT)^Δn with R = 0.08206 L atm mol^{-1} K^{-1}. (d) A mixture at 1000 K has P_SO_3 = 0.90 atm, P_SO_2 = 0.60 atm and P_O_2 = 0.45 atm. Calculate Qp, showing the substitution, and state the net direction.

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Section C: Exam Style

C1. The water gas shift reaction, $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$, has Kc = 4.0 at the experimental temperature. A vessel holds [CO] = 0.10 mol L^{-1}, [H_2O] = 0.10 mol L^{-1}, [CO_2] = 0.30 mol L^{-1} and [H_2] = 0.30 mol L^{-1}. (a) Write the expression for Qc. (b) Calculate Qc, showing the substitution. (c) State the net direction, and the change in each of the four concentrations as equilibrium is approached. (d) The vessel is suddenly expanded to twice its volume, so every concentration halves before any reaction occurs. Calculate the new value of Qc and state whether the net direction changes, supporting the answer with the coefficient totals on the two sides of the equation.

C2. Two reactions are given at the same temperature: (1) $C(s) + O_2(g) \rightleftharpoons CO_2(g)$, $K_1 = 2.0 \times 10^2$; (2) $2CO(g) + O_2(g) \rightleftharpoons 2CO_2(g)$, $K_2 = 4.0 \times 10^2$. The target reaction is $2C(s) + O_2(g) \rightleftharpoons 2CO(g)$. (a) State what must be done to reaction (1), write the manipulated reaction, and state its constant. (b) State what must be done to reaction (2), write the manipulated reaction, and state its constant. (c) Add the two manipulated reactions and show that the species that cancel leave exactly the target. (d) Calculate K for the target reaction.

C3. Judge each claim, supporting each verdict with the idea it depends on. (a) "In a Kc expression, each concentration is multiplied by its coefficient." (b) "Adding more of a solid reactant to an equilibrium mixture increases Q and shifts the reaction forward." (c) "Q = K means the reactant and product concentrations are equal." (d) "Adding argon to a gaseous equilibrium mixture at constant volume increases the total pressure, so Q changes and the system shifts." (e) "Writing a reaction backwards doubles its equilibrium constant."

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Answer Key

A1. (a) Kc = $[NOCl]^2$ / ($[NO]^2[Cl_2]$). (b) Kc = $[NO_2]^2$ / $[N_2O_4]$. (c) Products go on top and reactants underneath; every concentration is raised to the power of its own coefficient in the balanced equation, never multiplied by it. Only gases and aqueous species appear at all: anything written (s) or (l) for a pure solid or liquid is left out before the powers are even considered.

A2. (a) The concentration of a pure solid is fixed by its density and molar mass and does not change as the reaction proceeds, however much solid is present. That constant value is folded into K, so the solid is left out of the expression. (b) Kc = $[CO_2]$. (c) Kc = $[CO][H_2]$ / $[H_2O]$; the solid carbon is omitted.

A3. (a) The expressions are identical in form. Q is calculated from the concentrations at the moment in question, whatever they are; K is the value of the same expression when the mixture is at equilibrium at that temperature. (b) Q moves toward K: it rises toward K if it starts below K, falls toward K if it starts above K, and equals K once equilibrium is reached. (c) A change in temperature. Nothing else, not concentration, pressure, volume or a catalyst, changes K.

A4. (a) Net forward: product concentrations rise and reactant concentrations fall until Q climbs to K. (b) Net reverse: product concentrations fall and reactant concentrations rise until Q drops to K. (c) The mixture is at equilibrium: no net change, and every concentration stays constant even though both reactions continue.

A5. (a) The numerator holds one species squared, a power total of 2; the denominator holds one species squared and one to the first power, a power total of 3. Doubling every concentration multiplies the numerator by $2^2 = 4$ and the denominator by $2^3 = 8$, so Q is multiplied by 4/8 = ½: Q halves. (b) Numerator power total 2, denominator power total 1 + 1 = 2. Doubling multiplies the top by $2^2 = 4$ and the bottom by 2 × 2 = 4, so Q is multiplied by 4/4 = 1: Q is unchanged. (c) Whether the coefficient totals on the two sides are equal. Equal totals mean the scaling factor cancels between top and bottom; a larger reactant total, as in (a), makes Q shrink when concentrations rise.

B1. (a) Q = $[HI]^2$ / ($[H_2][I_2]$). (b) Q = $(1.00)^2$ / (0.125 × 0.125) = 1.00 / 0.015625 = 64. (c) Q = 64 = Kc, so this mixture is at equilibrium: no net change, and the concentrations stay at these values.

B2. (a) Q = $(0.80)^2$ / (0.20 × 0.20) = 0.64 / 0.040 = 16. (b) Q = 16 < Kc = 64, so the net reaction is forward. (c) [HI] rises; $[H_2]$ and $[I_2]$ fall, until Q has climbed to 64. The changes keep the coefficients' ratio: each 0.10 mol L^{-1} of HI formed consumes 0.050 mol L^{-1} of H_2 and 0.050 mol L^{-1} of I_2, so the two reactant concentrations fall together, by equal amounts, in this reaction.

B3. (a) Q = $(0.60)^2$ / (0.050 × 0.050) = 0.36 / 0.0025 = 144. (b) Q = 144 > Kc = 64, so the net reaction is reverse: HI is consumed net. (c) Q falls to 64. At that point the mixture is at equilibrium: the forward and reverse rates are equal, so the concentrations stop changing.

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B4. (a) Kc = 1/64 = 0.0156. Rule: reversing the equation inverts the constant. (b) Kc = √64 = 8. Rule: halving every coefficient takes the square root of the constant. (c) Kc = 64^2 = 4096. Rule: multiplying every coefficient by n raises the constant to the power n.

B5. (a) Q = $[CO_2]$: both solids are omitted. (b) Q = 0.020. Since 0.020 < Kc = 0.050, the net reaction is forward: $CaCO_3(s)$ decomposes, so the amount of $CaCO_3$ falls and $[CO_2]$ rises toward 0.050 mol L^{-1}. (c) Q = 0.085. Since 0.085 > 0.050, the net reaction is reverse: CO_2 reacts with CaO, so the amount of CaO(s) falls and $CaCO_3$ is formed. (d) Q stays 0.020. The added solid does not appear in the expression, and the concentration of a pure solid is the same however much of it is present, so neither Q nor the direction changes. (e) Both vessels settle at $[CO_2]$ = 0.050 mol L^{-1}, the value Kc fixes, because in each vessel Q moves until it equals Kc. The condition is that both solids are still present at that point: if every particle of $CaCO_3$ had decomposed, or every particle of CaO had reacted, the movement would stop short of 0.050 for lack of one reactant.

B6. (a) Kp = $(P_SO_3)^2$ / $((P_SO_2)^2 \times P_O_2)$. (b) Δn = 2 − (2 + 1) = −1. (c) Kp = Kc(RT)^Δn = 250 × $(0.08206 \times 1000)^{-1}$ = 250 / 82.06 = 3.0. Kp comes out far smaller than Kc here because Δn is negative: the reaction reduces the number of gas moles, so the conversion divides by RT rather than multiplying by it. (d) Qp = $(0.90)^2$ / $((0.60)^2 \times 0.45)$ = 0.81 / (0.36 × 0.45) = 0.81 / 0.162 = 5.0. Qp = 5.0 > Kp = 3.0, so the net reaction is reverse.

C1. (a) Qc = $[CO_2][H_2]$ / $([CO][H_2O])$. (b) Qc = (0.30 × 0.30) / (0.10 × 0.10) = 0.090 / 0.010 = 9.0. (c) Qc = 9.0 > Kc = 4.0, so the net reaction is reverse: $[CO_2]$ and $[H_2]$ fall while [CO] and $[H_2O]$ rise, until Qc falls to 4.0. (d) Halved concentrations give Qc = (0.15 × 0.15) / (0.05 × 0.05) = 0.0225 / 0.0025 = 9.0, unchanged. The coefficient totals are equal, 2 on each side, so halving multiplies the numerator and the denominator by the same factor, $(½)^2$. The direction is unchanged: still net reverse. Contrast A5(a), where the reactant side held the larger power total, 3 against 2, and the same scaling of every concentration did move Q.

C2. (a) Multiply reaction (1) by 2: $2C(s) + 2O_2(g) \rightleftharpoons 2CO_2(g)$, with constant $K_1^2 = (2.0 \times 10^2)^2 = 4.0 \times 10^4$. (b) Reverse reaction (2): $2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$, with constant $1/K_2 = 1 / (4.0 \times 10^2) = 2.5 \times 10^{-3}$. (c) Adding gives $2C(s) + 2O_2(g) + 2CO_2(g) \rightleftharpoons 2CO_2(g) + 2CO(g) + O_2(g)$. The $2CO_2$ appears on both sides and cancels, and one O_2 cancels from the two on the left, leaving $2C(s) + O_2(g) \rightleftharpoons 2CO(g)$, the target. (d) $K = K_1^2 \times (1/K_2) = 4.0 \times 10^4 / 4.0 \times 10^2 = 1.0 \times 10^2$.

C3. (a) False. Coefficients become powers, not multipliers: the SO_2 term in the sulfur trioxide expression is $[SO_2]^2$, and writing $2[SO_2]$ gives a different and wrong expression. (b) False. A pure solid does not appear in the expression at all, and its concentration cannot change, so adding more of it leaves Q exactly where it was and nothing shifts; B5(d) is this case with numbers. (c) False. Q = K is a statement about the ratio, not about equal concentrations: the B1 mixture, $[H_2]$ = $[I_2]$ = 0.125 mol L^{-1} with [HI] = 1.00 mol L^{-1}, has Q = Kc = 64 with concentrations far from equal. (d) False. At constant volume the argon changes no concentration and no partial pressure of any reacting species, so Q is unchanged, K is unchanged, and there is no shift; the total pressure rises,

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but no term in the expression contains it. (e) False. Reversing swaps products and reactants in the expression, so the new constant is the reciprocal, 1/K: the reversed hydrogen iodide reaction in B4 has Kc = 1/64 = 0.0156, not 128.

Continue Revising

Topic 7.3 Revision Notes

Topic 7.3 Quiz: Reaction Quotient and Equilibrium Constant

Previous Worksheet: Topic 7.2 Direction of Reversible Reactions

Next Worksheet: Topic 7.4 Calculating the Equilibrium Constant

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