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Topic 7.10: Reaction Quotient and Le Chatelier's Principle

Revision Notes · AP Chemistry

UNIT 7 · EQUILIBRIUM

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What Q Is

The reaction quotient Q has exactly the same form as the equilibrium constant K, but it uses the concentrations or partial pressures present right now.

For the general reaction

$aA + bB \rightleftharpoons cC + dD$

the reaction quotient is

$Q_c = [C]^c[D]^d \div [A]^a[B]^b$

Pure solids and pure liquids are left out, just as they are in K. If you are working with gases and pressures, Qp uses partial pressures in the same pattern.

Reading Q Against K

There are only three cases. If Q is less than K, the system is short on products, so the forward reaction runs and the mixture shifts right. If Q is greater than K, too much product is present, so the reverse reaction runs and the mixture shifts left. If Q equals K, the mixture is already at equilibrium and nothing shifts. This one comparison replaces all the memorized Le Chatelier rules.

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Concentration Stresses Change Q

Adding a reactant increases the denominator of Q, so Q falls below K and the reaction shifts right. Adding a product increases the numerator, so Q rises above K and the reaction shifts left. Removing a product lowers the numerator, so Q falls below K and the reaction shifts right.

Why K does not move here. Concentration changes never change K. Only temperature changes K. An exam favorite claims adding reactant increases Kc. It is wrong: K is constant at constant temperature, and the shift exists to bring Q back to the unchanged K. The shift only partly counteracts the disturbance: the new equilibrium holds more product than before, but some added reactant is still there at the end.

Pressure and Volume Stresses

These matter only for gases, and only when the two sides have different numbers of gas moles. Decreasing the volume raises every concentration at once, and whether Q rises or falls depends on the exponents.

Take this equilibrium, at equilibrium:

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

Halve the volume and every concentration doubles. The numerator of Q grows by a factor of $2^2 = 4$, while the denominator grows by $2^2 \times 2 = 8$. Overall Q becomes half its old value, so Q is now below Kc and the reaction shifts right, toward the side with fewer gas moles, 2 moles versus 3.

The shortcut: increasing pressure shifts equilibrium toward fewer gas moles; decreasing pressure shifts it toward more gas moles. When both sides have equal gas moles, a volume change scales the numerator and denominator identically, Q does not change, and nothing shifts. Dilution works the same way for aqueous equilibria: diluting to twice the volume halves every concentration, so a reaction with more dissolved particles on the product side shifts right.

Temperature Stresses Change K Itself

Temperature is the one stress that changes K rather than Q. Raising the temperature decreases K for an exothermic reaction and increases K for an endothermic one. Lowering the temperature does the reverse in each case.

The logic still runs through the Q versus K comparison. Suppose this exothermic equilibrium sits at equilibrium:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

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Raising the temperature drops K to a smaller value. At that instant Q is unchanged, so Q now exceeds the new smaller K and the system shifts left toward reactants until Q falls to the new K.

A common error is to say heating changes Q directly. It does not: heating changes K, and the mismatch between the old Q and the new K drives the shift. Always adjust K first, then compare Q with the new K.

Things That Do Nothing

A catalyst speeds up the forward and reverse reactions by exactly the same factor, so it gets you to equilibrium faster without moving it. Q, K, and all concentrations stay the same.

Adding an inert gas at constant volume changes the total pressure but not any concentration, so Q is untouched and nothing shifts.

Adding more of a pure solid or pure liquid also does nothing, because solids and liquids do not appear in Q. In this equilibrium, Q is simply $[CO_2]$:

$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

Connecting Q to Observations

The AP exam often ties the Q comparison to something you can see: the product here is deep red.

$Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{2+}(aq)$

Adding SCN^- enlarges the denominator of Q, so Q falls below K and the reaction shifts right, making more red complex and deepening the color. Removing Fe^{3+} would raise Q above K, shift the reaction left, and fade the color.

Worked Examples

Example 1: Q below K, shift right. The reaction

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

has Kc = 54. A vessel holds 1.0 M H_2, 1.0 M I_2, and 3.0 M HI. Then $Q = (3.0)^2 \div (1.0 \times 1.0) = 9.0$. Since 9.0 is less than 54, the forward reaction proceeds until Q climbs to 54.

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Example 2: Adding a reactant. The reaction

$2A(g) \rightleftharpoons B(g) + C(g)$

has Kc = 1.6. At equilibrium [A] = 0.50 M, [B] = 0.80 M, and [C] = 0.50 M. Then 0.30 mol/L of A is added, making [A] = 0.80 M. The new Q is $(0.80)(0.50) \div (0.80)^2 = 0.625$. Since 0.625 is below 1.6, the forward reaction runs to consume the added A until Q returns to 1.6.

Example 3: Halving the volume. The reaction

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

is at equilibrium, so Q = Kc. The volume is halved, doubling every concentration. The numerator of Q grows by $2^2 = 4$ and the denominator by $2^2 \times 2 = 8$, so the new Q is Kc ÷ 2, below Kc, and the reaction shifts right toward fewer gas moles until Q climbs back to Kc.

Example 4: Raising the temperature of an exothermic reaction. The reaction

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

is exothermic and sits at equilibrium, so Q = K. The temperature is raised, which lowers K for an exothermic reaction. At that instant Q is unchanged, so Q now exceeds the new, smaller K. The system shifts left toward reactants until Q falls to the new K, so the equilibrium amount of NH_3 at the higher temperature is lower than before heating.

Traps That Cost Marks in Topic 7.10

"Q less than K means shift left." Many students guess the opposite. Q less than K means the mixture is short on products, so the forward reaction runs and the mixture shifts right.

"Adding a reactant increases Kc." Kc is a constant at constant temperature. Adding a reactant changes Q, never K. The shift exists to bring Q back to the unchanged K.

"Heating changes Q directly." Heating changes K. Q only changes when a concentration or a partial pressure changes. For a temperature stress, adjust K first, then compare the unchanged Q with the new K.

"A catalyst shifts the equilibrium." A catalyst speeds both directions equally. It changes how fast equilibrium is reached and nothing else.

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"Adding an inert gas shifts the equilibrium because the pressure rises." At constant volume, an inert gas changes no concentration and no partial pressure of a reacting gas. Q is untouched, so nothing shifts.

"More solid reactant shifts the equilibrium right." Solids and liquids never appear in Q. Adding solid $CaCO_3$ to its equilibrium changes nothing at all.

Key Relationships to Remember

- Q has the same form as K but uses current concentrations or partial pressures; K is the target, Q is the snapshot
- Q less than K means shift right; Q greater than K means shift left; Q equal to K means equilibrium
- Adding a reactant or removing a product lowers Q below K and shifts the reaction right; adding a product or removing a reactant does the reverse
- Concentration, pressure, and volume stresses change Q; only temperature changes K
- Increasing pressure shifts equilibrium toward fewer gas moles; decreasing pressure shifts it toward more gas moles; equal gas moles means no shift
- For an exothermic reaction, raising the temperature lowers K; for an endothermic reaction, raising the temperature raises K
- A catalyst, an inert gas at constant volume, and added solid or liquid never shift an equilibrium
- Pure solids and pure liquids are omitted from Q, so Q for the $CaCO_3$ equilibrium is simply $[CO_2]$

Continue Revising

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Topic 7.10 Worksheet

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