

Unit 7: Equilibrium

Unit 7 is about chemical equilibrium. It covers what it means for forward and reverse reactions to balance, how to write and use the equilibrium constant, how to predict which way a reaction will shift, and how the same ideas apply to solids dissolving in water.

AP Chemistry

Equilibrium

About 12 minutes to read

How to use this guide

Read it in order the first time because the topics build on each other. Dynamic equilibrium motivates the equilibrium constant, the constant lets you write the Q expression, comparing Q to K predicts the direction of shift, and Le Châtelier's principle is the same comparison stated as a rule. Solubility equilibrium at the end reuses every idea from the first half with solids.

After the first read, use the trap boxes and the confusion table to review the distinctions that exam questions test most often. Finish with the practice questions, then complete the recall check on the last page out loud and note any items you cannot explain yet.

What this unit is worth. Equilibrium is about 7 to 9 percent of the AP Chemistry exam. It also carries more weight than that number suggests, because Unit 8 (acids and bases) is equilibrium applied to a special case. If the Q -versus- K logic feels shaky now, acid-base equilibrium will feel shaky later.

7.1 Introduction to Equilibrium

A **reversible process** is one that can proceed in both the forward and reverse directions. Many chemical reactions are reversible: products can react to re-form reactants, so the reaction does not have to run only one way.

Dynamic equilibrium is the state in which the forward and reverse rates of a process are equal. Because the two rates match, there is no net change, and concentrations remain constant. The word dynamic matters. Molecules keep reacting in both directions. What stops changing is the overall amounts, not the activity.

Trap. Equilibrium does not mean the concentrations of reactants and products are equal. It means their ratio has settled at the value the equilibrium constant demands. Equal concentrations happen only for special K values near 1.

7.2 Direction of Reversible Reactions

At any moment, the **net direction of a reversible reaction** is set by the faster of the two reactions. If the forward reaction is running faster than the reverse, there is net conversion of reactants to products. If the reverse is faster, there is net conversion of products to reactants.

When the two rates become equal, the system is at equilibrium and net conversion stops. This is why equilibrium concentrations stay constant: every molecule of product being formed is matched by one being consumed. The rates equalize because the forward rate depends on reactant concentrations and the reverse rate depends on product concentrations, so as reactants are used up and products build up, the two rates converge.

Trap. "No net change" is not the same as "nothing happening." A question that describes constant concentrations over time is describing dynamic equilibrium, with both reactions still running. Only the net change is zero.

7.3 Reaction Quotient and Equilibrium Constant

The **reaction quotient (Q)** has the same form as the equilibrium constant expression, but it is calculated with current concentrations or partial pressures, not necessarily equilibrium ones. Q_c uses concentrations and Q_p uses partial pressures. Q is a snapshot of where the mixture stands right now.

The **equilibrium constant (K)** is the value of the reaction quotient when the system is at equilibrium. K_c uses equilibrium concentrations and K_p uses equilibrium partial pressures. Pure solids and pure liquids are omitted from the expression, because their concentrations do not change as the reaction proceeds.

The **law of mass action** gives the form. For $aA + bB \rightleftharpoons cC + dD$, $K = [C]^c[D]^d / ([A]^a[B]^b)$. Products go on top, reactants on the bottom, and each coefficient becomes an exponent. For $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, $K_c = [NH_3]^2 / ([N_2][H_2]^3)$.

Trap. The two most common expression errors are forgetting to raise concentrations to the power of their coefficients and including pure solids or liquids. For $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$, the numerator is $[SO_3]^2$, not $[SO_3]$. For $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$, the expression is just $K = [CO_2]$, since both solids are omitted.

7.4 Calculating the Equilibrium Constant

Determining K from measurements is how the constant is found experimentally. Measure the equilibrium concentrations or partial pressures, then substitute them into the mass-action expression.

For $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, suppose measurements at equilibrium give $[H_2] = 0.10$ M, $[I_2] = 0.10$ M, and $[HI] = 0.74$ M. Then $K_c = [HI]^2 / ([H_2][I_2]) = (0.74)^2 / (0.10 \times 0.10) = 0.5476 / 0.010 = 54.8$, about 55. This is the real value for this reaction near 425 °C, and it is the kind of arithmetic the exam expects you to do cleanly.

Trap. K must be calculated from equilibrium concentrations, not initial ones. If a question gives you starting amounts, you need the change to equilibrium first, usually through an ICE table, before you touch the K expression.

7.5 Magnitude of the Equilibrium Constant

The **magnitude of K** tells you where equilibrium lies. A large K ($K \gg 1$) means the reaction proceeds essentially to completion and favors products at equilibrium. A small K ($K \ll 1$) means the reaction barely proceeds and favors reactants. A K near 1 means appreciable amounts of both reactants and products are present at equilibrium.

For the Haber process, $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$, K_c is very large at room temperature, so the equilibrium mixture is mostly ammonia. At the high temperatures used industrially, K is much smaller, which is one reason the process needs high pressure to get a useful yield.

Trap. A large K says nothing about speed. K describes the position of equilibrium, which is thermodynamics. How fast the reaction gets there is kinetics, governed by activation energy. A reaction can have an enormous K and still be immeasurably slow at room temperature.

7.6 Properties of the Equilibrium Constant

Manipulating equilibrium constants follows three rules. Reversing a reaction gives $1/K$. Multiplying all coefficients by a number c gives K^c . Adding two reactions multiplies their K values to get the K of the combined reaction.

Manipulating the reaction quotient uses the same rules, because K and Q have identical mathematical forms. Reversing a reaction inverts Q, scaling coefficients by c raises Q to the power c , and adding reactions multiplies their Q values.

Example: if $\text{A(g)} \rightleftharpoons \text{B(g)}$ has $K_c = 4.0$, then the reverse reaction $\text{B(g)} \rightleftharpoons \text{A(g)}$ has $K_c = 1/4.0 = 0.25$, and the doubled reaction $2\text{A(g)} \rightleftharpoons 2\text{B(g)}$ has $K_c = (4.0)^2 = 16$. If a question asks for $2\text{B(g)} \rightleftharpoons 2\text{A(g)}$, apply both steps: reverse and double, giving $(1/4.0)^2 = 1/16 = 0.0625$.

Trap. When a reaction is both reversed and scaled, students often apply only one operation. Write each manipulation as its own step. Reverse first, then scale, or scale first, then reverse. The result is the same either way, but skipping a step is where the points go.

7.7 Calculating Equilibrium Concentrations

Predicting direction with Q vs. K is the central comparison of the unit. If $Q < K$, the reaction shifts toward products. If $Q > K$, it shifts toward reactants. If $Q = K$, the system is already at equilibrium and nothing shifts.

The logic is straightforward. $Q < K$ means the current ratio of products to reactants is below the equilibrium ratio, so the system makes more products to catch up. $Q > K$ means the ratio is above equilibrium, so the system converts products back to reactants.

Worked example: for $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$ with $K_c = 55$, a mixture has $[\text{H}_2] = 0.20 \text{ M}$, $[\text{I}_2] = 0.20 \text{ M}$, and $[\text{HI}] = 0.10 \text{ M}$. $Q_c = (0.10)^2 / (0.20 \times 0.20) = 0.010 / 0.040 = 0.25$. Since $0.25 < 55$, the reaction shifts toward products, making more HI.

To find actual equilibrium concentrations from starting amounts, set up an ICE table: list **I**nitial concentrations, the **C**hange in terms of x , and the **E**quilibrium expressions, then substitute the equilibrium row into the K expression and solve for x .

Trap. Q and K look identical but answer different questions. Q describes the mixture right now. K describes the mixture at equilibrium. A question that gives you concentrations and asks for the direction of shift wants Q compared to K , not K recalculated.

7.8 Representations of Equilibrium

Particulate representations of equilibrium are particle-level diagrams of an equilibrium mixture. They show constant macroscopic concentrations arising from continuous forward and reverse reactions at the molecular level.

In such a diagram, the number of each type of particle stays the same from one frame to the next, on average, even though individual molecules keep reacting. If a diagram shows A particles decreasing and B particles increasing over time, the system is not yet at equilibrium. It is still shifting. At equilibrium the counts hold steady.

7.9 Introduction to Le Châtelier's Principle

Le Châtelier's principle says that when a system at equilibrium is disturbed, it shifts in the direction that restores equilibrium. The common disturbances are changes in concentration, changes in pressure or volume for gases, and changes in temperature.

Adding a reactant shifts the reaction toward products to consume the added material. Removing a product does the same thing, which is why industrial processes often remove product continuously. Increasing the pressure by decreasing the volume shifts the equilibrium toward the side with fewer moles of gas. For temperature, treat heat as a reactant or product: in an exothermic reaction, heat is a product, so raising the temperature shifts the equilibrium toward reactants.

Trap. A catalyst does not shift equilibrium. It lowers the activation energy for both the forward and reverse reactions equally, so the system reaches equilibrium faster but the equilibrium position and K are unchanged.

7.10 Reaction Quotient and Le Châtelier's Principle

Q vs. K after a disturbance is Le Châtelier's principle restated in the language of topic 7.7. A concentration disturbance changes Q but not K . A temperature change changes K itself. After either disturbance, the system shifts until $Q = K$ again.

Adding more reactant makes the denominator of Q larger, so Q drops below K , and the system shifts toward products to restore the ratio. Raising the temperature of an exothermic reaction lowers K , so the equilibrium ratio of products to reactants falls, and the system shifts toward reactants. This is the same prediction the principle gives, now with the mechanism visible.

Trap. Students often say a temperature change "shifts Q ." It does not. Temperature is the one disturbance that changes K itself, because K depends on temperature. Concentration and pressure changes move Q ; temperature moves K . Keep which one changes straight.

7.11 Introduction to Solubility Equilibria

The **solubility-product constant (K_{sp})** is the equilibrium constant for a slightly soluble salt dissolving. For $M_aX_b(s) \rightleftharpoons aM^+ + bX^-$, $K_{sp} = [M^+]^a[X^-]^b$. The solid is omitted from the expression, as with all pure solids. A larger K_{sp} means a more soluble salt.

Molar solubility is the moles of salt that dissolve per liter of saturated solution. It can be calculated from K_{sp} , and K_{sp} can be calculated from it. For $AgCl(s) \rightleftharpoons Ag^+ + Cl^-$ with $K_{sp} = 1.8 \times 10^{-10}$: if s mol/L dissolves, then $[Ag^+] = [Cl^-] = s$, so $K_{sp} = s^2$ and $s = \sqrt{1.8 \times 10^{-10}} \approx 1.3 \times 10^{-5}$ M.

On **K_{sp} and solubility**: the solubility rules can be related quantitatively to K_{sp} . Salts with K_{sp} values greater than 1 are considered soluble. Salts far below 1 are the slightly soluble ones whose equilibria this topic treats.

Trap. The exponents in the K_{sp} expression come from the coefficients, and forgetting them is the classic error. For $PbI_2(s) \rightleftharpoons Pb^{2+} + 2I^-$, $K_{sp} = [Pb^{2+}][I^-]^2 = (s)(2s)^2 = 4s^3$, not s^2 . The 2 on the iodide appears twice: once in the concentration $2s$ and once as the exponent.

7.12 Common-Ion Effect

The **common-ion effect** is the decrease in a salt's solubility when the solution already contains one of its ions. The added ion shifts the dissolution equilibrium toward the solid, which is Le Châtelier's principle applied to a solubility equilibrium.

Silver chloride dissolves less in 0.10 M NaCl than in pure water. In pure water, $s \approx 1.3 \times 10^{-5}$ M. In 0.10 M Cl^- , $K_{sp} = [Ag^+][Cl^-] = (s)(0.10 + s) \approx (s)(0.10)$, so $s \approx 1.8 \times 10^{-10} / 0.10 = 1.8 \times 10^{-9}$ M. The common ion suppresses the solubility by about four orders of magnitude. The s in $0.10 + s$ is dropped because it is negligible next to 0.10, a standard approximation.

Trap. The common ion must actually be one of the salt's ions. Adding $NaNO_3$ to $AgCl$ does nothing to the solubility, because neither Na^+ nor NO_3^- appears in the dissolution equilibrium. Check the formulas before predicting an effect.

Confusions That Cost Points

PAIR	HOW TO KEEP THEM STRAIGHT
Q vs K	Q uses current concentrations and tells you where the mixture is now. K uses equilibrium concentrations and tells you where it is headed. Compare them to find the direction of shift.
Large K vs fast reaction	K is about position, from thermodynamics. Speed is about activation energy, from kinetics. A huge K can belong to a very slow reaction.
Concentration change vs temperature change	Changing concentration or pressure moves Q. Changing temperature moves K itself. Only temperature changes the constant.
Equilibrium vs equal concentrations	Equilibrium means the ratio [products]/[reactants] has settled at K. The concentrations are equal only when K happens to be near 1.
Catalyst vs shift	A catalyst speeds up both directions equally. It gets you to equilibrium faster but never changes the position or K.
K_{sp} exponents	The ion coefficients become exponents in the K_{sp} expression and multiply the molar solubility. For a 1:2 salt, $K_{sp} = 4s^3$.
Shifting toward products vs K increasing	A shift toward products after adding reactant does not change K. K changes only with temperature. The shift restores Q to the same K.
Omitting solids and liquids	Pure solids and pure liquids never appear in K, Q, or K_{sp} expressions. Adding more solid reactant or product does not shift the equilibrium.

Practice Questions

Original questions written for this guide in the style of the AP exam. Answers and explanations are on the next page, so complete the questions before checking them.

1. For the reaction $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$, $K_c = 2.8 \times 10^2$ at a certain temperature. A mixture contains $[\text{SO}_2] = 0.10 \text{ M}$, $[\text{O}_2] = 0.10 \text{ M}$, and $[\text{SO}_3] = 2.0 \text{ M}$. In which direction will the reaction proceed?

- A. Toward products
- B. Toward reactants
- C. It is at equilibrium; no net shift
- D. Cannot be determined from this information

2. For the reaction $\text{A}(\text{g}) \rightleftharpoons \text{B}(\text{g})$, $K_c = 3.0$. What is K_c for the reaction $2\text{B}(\text{g}) \rightleftharpoons 2\text{A}(\text{g})$?

- A. 6.0
- B. $1/3$
- C. $1/9$
- D. 9.0

3. The reaction $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ is exothermic. Which change increases the equilibrium yield of NH_3 ?

- A. Increasing the temperature
- B. Decreasing the pressure
- C. Removing NH_3 as it forms
- D. Adding a catalyst

4. The K_{sp} of PbI_2 is 9.8×10^{-9} . What is the molar solubility of PbI_2 in pure water?

- A. $9.9 \times 10^{-5} \text{ M}$
- B. $1.3 \times 10^{-3} \text{ M}$
- C. $9.8 \times 10^{-9} \text{ M}$
- D. $2.1 \times 10^{-3} \text{ M}$

Answer Key

- 1. B.** $Q_c = [\text{SO}_3]^2 / ([\text{SO}_2]^2[\text{O}_2]) = (2.0)^2 / ((0.10)^2(0.10)) = 4.0 / 0.0010 = 4000$. Since Q (4000) > K (280), the reaction shifts toward reactants. A reverses the comparison. C would require $Q = K$. D is wrong because Q and K are both known.
- 2. C.** Reversing $A \rightleftharpoons B$ gives $K = 1/3.0$ for $B \rightleftharpoons A$. Doubling the coefficients squares it: $(1/3.0)^2 = 1/9$. A doubles K instead of applying the rules. B applies the reversal but forgets to square for the doubled coefficients. D squares without reversing.
- 3. C.** Removing product lowers Q below K , so the system shifts toward products to restore equilibrium, making more NH_3 . A is wrong because the reaction is exothermic, so higher temperature favors reactants. B is wrong because lower pressure favors the side with more moles of gas, which is the reactant side (4 moles vs 2). D is wrong because a catalyst never changes the equilibrium position.
- 4. B.** $\text{PbI}_2(s) \rightleftharpoons \text{Pb}^{2+} + 2\text{I}^-$, so $K_{\text{sp}} = [\text{Pb}^{2+}][\text{I}^-]^2 = (s)(2s)^2 = 4s^3$. Then $s = \sqrt[3]{(9.8 \times 10^{-9} / 4)} = \sqrt[3]{(2.45 \times 10^{-9})} \approx 1.3 \times 10^{-3} \text{ M}$. A comes from forgetting the coefficient on iodide and solving $s = \sqrt{K_{\text{sp}}}$. C confuses K_{sp} with the solubility itself. D comes from an arithmetic slip in the cube root.

One-Page Recall Check

- Define a reversible process and explain what makes equilibrium dynamic rather than static.
- Explain how the relative rates of the forward and reverse reactions determine the net direction.
- State the difference between Q and K , and when each one uses current vs equilibrium concentrations.
- Write the law of mass action for $aA + bB \rightleftharpoons cC + dD$ and explain why pure solids and liquids are omitted.
- Calculate K_c from a set of measured equilibrium concentrations.
- Explain what $K \gg 1$ and $K \ll 1$ each imply, and why K says nothing about reaction speed.
- State the three rules for manipulating K (and Q): reversing, scaling, and adding reactions.
- Use Q vs K to predict the direction of shift, and explain the logic behind each case.
- Describe how to set up an ICE table to find equilibrium concentrations.
- Explain what a particulate diagram shows at equilibrium vs before equilibrium is reached.
- State Le Châtelier's principle and predict shifts for concentration, pressure, and temperature changes.
- Explain why a concentration change alters Q while a temperature change alters K .
- Write the K_{sp} expression for M_aX_b and calculate molar solubility from it.
- Explain the common-ion effect with the $AgCl$ in $NaCl$ example.

Study this unit in Rycal. Open the Equilibrium deck under AP Chemistry at rycal.web.app/apchem. The deck covers the terms in this guide, and its practice questions target the same traps named here. If you have a test date, add it in the Test Planner. You can also start your next review with a Brain Dump, then check what you missed against this guide.

Key terms for this unit

Reversible process, Dynamic equilibrium, Net direction of a reversible reaction, Reaction quotient (Q), Equilibrium constant (K), Law of mass action, Determining K from measurements, Magnitude of K , Manipulating equilibrium constants, Manipulating the reaction quotient (Q), Predicting direction with Q vs. K , Particulate representations of equilibrium, Le Châtelier's principle, Q vs. K after a disturbance, Solubility-product constant (K_{sp}), Molar solubility, K_{sp} and solubility, Common-ion effect.

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