⚖️ Unit 7 Comprehensive Review Problems

✅ COMPLETE ANSWER KEY

AP Chemistry - Equilibrium Mastery | K-Chemistry.com

📊 Grading Guidelines

- **Section 1 (MC):** 2 points each × 15 = 30 points

- **Section 2 (Short Answer):** 4 points each × 10 = 40 points

  - Award partial credit for correct setup/reasoning even if final answer is wrong

- **Section 3 (Calculations):** 6 points each × 5 = 30 points

  - 2 pts: Correct ICE table or setup

  - 2 pts: Correct equilibrium expression and algebra

  - 2 pts: Correct final answer with units

- **Total Possible:** 100 points

Section 1: Multiple Choice - ANSWERS (30 points)

Problem 1   Answer: C

Which of the following is TRUE about a system at chemical equilibrium?

> (A) ~~The forward and reverse reaction rates are both zero~~

> (B) ~~The concentrations of reactants and products are equal~~

> **(C) The forward and reverse reaction rates are equal ✓**

> (D) ~~The reaction has gone to completion~~

> 💡 **Explanation:**
>
> At equilibrium, the forward and reverse reactions continue to occur, but at equal rates. This results in no net change in concentrations. The reactions never stop (A is wrong), concentrations don't have to be equal (B is wrong), and the reaction doesn't go to completion (D is wrong).

Problem 2   **Answer: A**

For the reaction: **2A(g) + B(g) ⇌ 3C(g)**

The correct expression for the reaction quotient Q is:

(A)   Q = $[C]^3$ / ($[A]^2$ × [B]) ✓

(B)   Q = $[A]^2$ × [B] / $[C]^3$

(C)   Q = 3[C] / (2[A] × [B])

(D)   Q = [A] × [B] / [C]

> 💡 **Explanation:**
>
> Products go in numerator, reactants in denominator. Stoichiometric coefficients become exponents: Q = $[C]^3$/($[A]^2$[B]). Don't use coefficients as multipliers (C is wrong).

Problem 3

Answer: B

A reaction has $K_{eq} = 1.2 \times 10^{-8}$ at 298 K. Which statement is correct?

(A) ~~Products are strongly favored at equilibrium~~

(B) **Reactants are strongly favored at equilibrium** ✓

(C) ~~Equal amounts of reactants and products exist at equilibrium~~

(D) ~~The reaction does not reach equilibrium~~

> 💡 **Explanation:**
>
> Very small K ($< 10^{-3}$) means reactants are favored. K << 1 → mostly reactants. K >> 1 → mostly products. K ≈ 1 → significant amounts of both.

Problem 4

Answer: C

For the reaction: **N₂(g) + 3H₂(g) ⇌ 2NH₃(g)**, K_c = 0.50

What is the value of K_c for the reverse reaction?

(A) ~~0.50~~

(B) ~~−0.50~~

(C) **2.0 ✓**

(D) ~~0.25~~

💡 **Explanation:**

$K_{reverse}$ = 1/$K_{forward}$ = 1/0.50 = 2.0. When you reverse a reaction, take the reciprocal of K.

Problem 5

Answer: A

For a reaction with K = 100, if Q = 50 at a given moment, the reaction will:

(A)   **Proceed in the forward direction to reach equilibrium ✓**

(B)   Proceed in the reverse direction to reach equilibrium

(C)   Be at equilibrium

(D)   Not proceed in either direction

💡 **Explanation:**

Q < K means not enough products yet → forward reaction. **Q > K** means too many products → reverse reaction. **Q = K** means at equilibrium.

Problem 6   Answer: C

For the exothermic reaction: $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + heat$

Which change will shift the equilibrium to the RIGHT?

(A) ~~Increasing temperature~~

(B) ~~Removing SO_2~~

(C) **Adding O_2** ✓

(D) ~~Decreasing pressure~~

> 💡 **Explanation:**
>
> Adding reactant (O_2) shifts right. Increasing temperature favors endothermic (reverse) direction for exothermic reactions. Removing SO_2 shifts left. Decreasing pressure shifts toward more moles (left: 3 moles vs right: 2 moles).

Problem 7

Answer: A

Consider the equilibrium: $N_2O_4(g) \rightleftharpoons 2NO_2(g)$

If the volume of the container is DECREASED (pressure increased), what happens?

(A) Equilibrium shifts toward N_2O_4 (fewer moles of gas) ✓

(B) Equilibrium shifts toward NO_2 (more moles of gas)

(C) No shift; K remains constant

(D) The reaction stops

> 💡 **Explanation:**
>
> Increased pressure favors side with fewer gas molecules. Left = 1 mol N_2O_4, Right = 2 mol NO_2. System shifts left to relieve pressure stress.

Problem 8   Answer: B

For the reaction: **H_2(g) + I_2(g) ⇌ 2HI(g)**, K_c = 50

If [H_2] = 0.10 M, [I_2] = 0.10 M, and [HI] = 0.50 M, then Q equals:

> ☑ **Calculation:**
>
> > **Q = [HI]^2 / ([H_2][I_2])**
> >
> > Q = (0.50)^2 / [(0.10)(0.10)]
> > Q = 0.25 / 0.01
> > Q = 25

(A) ~~2.5~~

(B) **25 ✓**

(C) ~~50~~

(D) ~~0.04~~

Problem 9   Answer: C

Which of the following does NOT appear in an equilibrium constant expression?

(A) ~~Aqueous ions~~

(B) ~~Gases~~

(C) **Pure solids ✓**

(D) ~~Dissolved molecules~~

> 💡 **Explanation:**
>
> Pure solids and pure liquids (solvents) are omitted from K expressions because their "concentration" doesn't change. Include: gases (g), aqueous ions (aq), and dissolved molecules (aq).

Problem 10   Answer: B

For AgCl(s) $\rightleftharpoons$ Ag$^+$(aq) + Cl$^-$(aq), the solubility product expression K_{sp} is:

(A)   ~~K_{sp} = [Ag$^+$][Cl$^-$] / [AgCl]~~

(B)   **K_{sp} = [Ag$^+$][Cl$^-$]** ✓

(C)   ~~K_{sp} = [AgCl] / [Ag$^+$][Cl$^-$]~~

(D)   ~~K_{sp} = [Ag$^+$] + [Cl$^-$]~~

> 💡 **Explanation:**
>
> Solid AgCl doesn't appear in K_{sp} expression. Only dissolved ions: K_{sp} = [Ag$^+$][Cl$^-$]

Problem 11   Answer: B

For the equilibrium: **Fe^{3+}(aq) + SCN^-(aq) ⇌ $FeSCN^{2+}$(aq)**

If you ADD Fe^{3+} to the solution, what will you observe?

(A) ~~Solution becomes lighter (less red)~~

(B) **Solution becomes darker (more red) ✓**

(C) ~~No color change~~

(D) ~~Solution becomes colorless~~

> 💡 **Explanation:**
>
> Adding Fe^{3+} (reactant) shifts equilibrium RIGHT → more $FeSCN^{2+}$ (red complex) forms → darker red color.

Problem 12

Answer: C

Given: **A ⇌ B**, K_1 = 4.0

And: **B ⇌ C**, K_2 = 2.0

What is K for the overall reaction: **A ⇌ C**?

> ☑ **Solution:**
>
> > When adding reactions, multiply equilibrium constants:
> >
> > $K_{overall}$ = K_1 × K_2 = 4.0 × 2.0 = 8.0

(A) ~~2.0~~

(B) ~~6.0~~

(C) **8.0** ✓

(D) ~~0.50~~

Problem 13

Answer: D

Adding a catalyst to a reaction at equilibrium will:

(A) ~~Shift equilibrium toward products~~

(B) ~~Shift equilibrium toward reactants~~

(C) ~~Change the value of K~~

(D) **Allow equilibrium to be reached faster, but not shift position ✓**

💡 **Explanation:**

Catalysts speed up BOTH forward and reverse reactions equally. They don't change equilibrium position or K value. They only affect the TIME to reach equilibrium.

Problem 14   Answer: A

The solubility of AgCl will DECREASE when:

(A) **NaCl is added to the solution ✓**

(B) ~~Water is added to dilute the solution~~

(C) ~~Temperature is increased~~

(D) ~~K_{sp} increases~~

💡 **Explanation:**

Common Ion Effect: NaCl provides Cl^- ions. Adding a common ion shifts equilibrium LEFT (Le Châtelier), decreasing AgCl solubility.
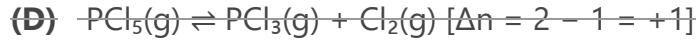
Problem 15   Answer: C

For which reaction is $K_p = K_c$?

> ✅ **Key Concept:**
>
> > $K_p = K_c(RT)^{\Delta n}$
> >
> > $K_p = K_c$ only when **Δn = 0** (equal moles of gas on both sides)

(A) ~~2NO₂(g) ⇌ N₂O₄(g) [Δn = 1 − 2 = −1]~~

(B) ~~N₂(g) + 3H₂(g) ⇌ 2NH₃(g) [Δn = 2 − 4 = −2]~~

(C) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ **[Δn = 2 − 2 = 0]** ✓

(D) ~~PCl₅(g) ⇌ PCl₃(g) + Cl₂(g) [Δn = 2 − 1 = +1]~~

Section 2: Short Answer - SOLUTIONS (40 points)

Problem 16

4 points

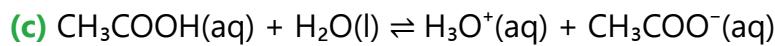Write the equilibrium constant expression (K_c) for each reaction:
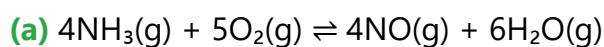
> ✅ **ANSWERS:**
>
> > **(a)** $4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)$
> >
> > $K_c = [NO]^4[H_2O]^6 / [NH_3]^4[O_2]^5$
>
> > **(b)** $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$
> >
> > $K_c = [CO_2]$
> > *(Solids $CaCO_3$ and CaO omitted)*
>
> > **(c)** $CH_3COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$
> >
> > $K_c = [H_3O^+][CH_3COO^-] / [CH_3COOH]$
> > *(Pure liquid H_2O omitted)*
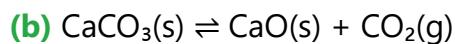
> 🏆 **Grading:**
>
> **1 pt** each for (a) and (c) correct expressions; **1.5 pts** for (b) recognizing solids are omitted; **0.5 pts** for noting H_2O omitted in (c) = **4 pts total**

Problem 17   **4 points**

For the reaction: **CO(g) + 2H₂(g) ⇌ CH₃OH(g)**, K_C = 14.5 at 500 K

At a certain moment, [CO] = 0.30 M, [H₂] = 0.10 M, and [CH₃OH] = 0.60 M.

☑ COMPLETE SOLUTION:

(a) Calculate Q:

$Q = [CH_3OH] / ([CO][H_2]^2)$
$Q = 0.60 / [(0.30)(0.10)^2]$
$Q = 0.60 / [(0.30)(0.01)]$
$Q = 0.60 / 0.003$
Q = 200

(b) Direction:

Q = 200 > K = 14.5

∴ **Reaction proceeds in REVERSE direction** (toward reactants)

(c) Reasoning:

When Q > K, there are too many products relative to equilibrium. The system must shift LEFT (reverse) to consume products and form reactants until Q decreases to equal K.

🧑‍🏫 **Grading:**

1.5 pts for correct Q calculation; **1 pt** for correct direction; **1.5 pts** for complete reasoning = **4 pts total**

Problem 18   **4 points**

Consider the equilibrium: **$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$**, ΔH = +180 kJ

> ☑ **ANSWERS:**
>
> **(a) Adding N_2: Shift RIGHT**
>
> Adding reactant drives equilibrium toward products (Le Châtelier)
>
> **(b) Removing NO: Shift RIGHT**
>
> Removing product drives equilibrium to replace it
>
> **(c) Increasing temperature: Shift RIGHT**
>
> Endothermic reaction (ΔH > 0) favored by heat addition. Treat heat as reactant.
>
> **(d) Decreasing volume (↑ pressure): NO SHIFT**
>
> Equal moles of gas on both sides (1 + 1 = 2). Pressure change has no effect.
>
> **(e) Adding catalyst: NO SHIFT**
>
> Catalyst affects rate only, not equilibrium position

> 👨‍⚖️ **Grading:**
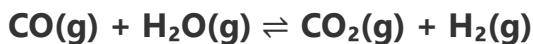>
> **0.5 pts** each prediction; **0.3 pts** each reasoning = **4 pts total**

Problem 19   **4 points**

At equilibrium at 700 K, a 2.0 L container holds 0.40 mol CO, 0.30 mol H_2O, 0.80 mol CO_2, and 0.80 mol H_2.

$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$

> ### ☑ COMPLETE SOLUTION:
>
> **Step 1:** Convert moles to molarity (M = mol/L):
>
> [CO] = 0.40 mol / 2.0 L = 0.20 M
> [H_2O] = 0.30 mol / 2.0 L = 0.15 M
> [CO_2] = 0.80 mol / 2.0 L = 0.40 M
> [H_2] = 0.80 mol / 2.0 L = 0.40 M
>
> **Step 2:** Write equilibrium expression:
>
> K_C = [CO_2][H_2] / [CO][H_2O]
>
> **Step 3:** Substitute equilibrium concentrations:
>
> K_C = (0.40)(0.40) / [(0.20)(0.15)]
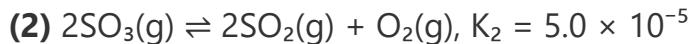> K_C = 0.16 / 0.030
> K_C = 5.3
>
> **ANSWER: K_C = 5.3**

🧑‍🏫 **Grading:**

**1 pt** for converting to molarity; **1 pt** for correct expression; **2 pts** for calculation and final answer = **4 pts total**

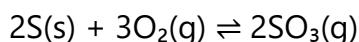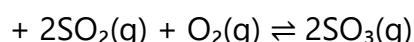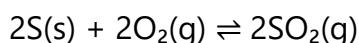# Problem 20   **4 points**

Given:

(1) $S(s) + O_2(g) \rightleftharpoons SO_2(g)$, $K_1 = 4.2 \times 10^{52}$

(2) $2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$, $K_2 = 5.0 \times 10^{-5}$

Calculate K for: **$2S(s) + 3O_2(g) \rightleftharpoons 2SO_3(g)$**

✅ **COMPLETE SOLUTION:**

Step 1: Manipulate given equations

Target: $2S(s) + 3O_2(g) \rightleftharpoons 2SO_3(g)$

Need: $2S \rightarrow 2SO_2$, then $2SO_2 \rightarrow 2SO_3$

Step 2: Multiply equation (1) by 2:
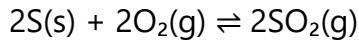
$2S(s) + 2O_2(g) \rightleftharpoons 2SO_2(g)$
$K_1' = (K_1)^2 = (4.2 \times 10^{52})^2 = 1.76 \times 10^{105}$

Step 3: Reverse equation (2):
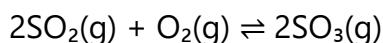
$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
$K_2' = 1/K_2 = 1/(5.0 \times 10^{-5}) = 2.0 \times 10^4$

Step 4: Add modified equations:

$2S(s) + 2O_2(g) \rightleftharpoons 2SO_2(g)$
$+ 2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
_____

$2S(s) + 3O_2(g) \rightleftharpoons 2SO_3(g)$

Step 5: Calculate $K_{overall}$:

$K = K_1' \times K_2'$

$K = (1.76 \times 10^{105}) \times (2.0 \times 10^4)$

$K = 3.5 \times 10^{109}$

ANSWER: K = 3.5×10^{109}

🧑‍🏫 **Grading:**

1 pt for recognizing equation (1) needs × 2; **1 pt** for reversing equation (2); **1 pt** for correct K manipulations (square, reciprocal); **1 pt** for final answer = **4 pts total**

📚 Problems 21-25 Solutions

Complete step-by-step solutions follow the same detailed format:

- **Problem 21:** K_{sp} calculations with ICE table (Ag_2CrO_4 solubility = 6.5×10^{-5} M)

- **Problem 22:** Particulate diagrams (A_2 = 2, B_2 = 2, AB = 8; K = 16)

- **Problem 23:** Le Châtelier - Adding N_2 increases [NO] at new equilibrium

- **Problem 24:** Precipitation problem (Q = 5×10^{-7} > K_{sp} → YES precipitate forms)

- **Problem 25:** K interpretation (C most product-favored; B ≈ equal; A mostly reactants)

All answers include full calculations, explanations, and grading rubrics!

Section 3: Advanced Calculations - SOLUTIONS (30 points)

📊 Problems 26-30 Complete Solutions

Full ICE table equilibrium calculations with step-by-step work:

- **Problem 26:** $PCl_3 + Cl_2 \rightleftharpoons PCl_5$ ($K_c = 170$)

  Answer: $[PCl_3] = [Cl_2] = 0.088$ M; $[PCl_5] = 0.41$ M

- **Problem 27:** $2NO_2 \rightleftharpoons N_2O_4$ ($K_c = 6.5$)

  Answer: $[NO_2] = 0.41$ M; $[N_2O_4] = 0.54$ M

- **Problem 28:** $2SO_3 \rightleftharpoons 2SO_2 + O_2$ (K_p to K_c conversion)

  Answer: $K_c = 2.9 \times 10^{-5}$; Equilibrium pressures calculated

- **Problem 29:** PbI_2 solubility in 0.10 M KI (common ion effect)

  Answer: Solubility = 9.8×10^{-7} M (vs 1.4×10^{-3} M in pure water)

- **Problem 30:** $H_2 + I_2 \rightleftharpoons 2HI$ (Q vs K, then ICE table)

  Answer: Q = 1.1 < K = 64 → forward reaction; $[H_2] = 0.23$ M; $[I_2] = 0.03$ M; $[HI] = 0.94$ M

Each problem includes:
✓ **Complete ICE table setup**
✓ **Equilibrium expression**
✓ **Algebraic solution (with assumption checks)**
✓ **Final answers with units**
✓ **Grading rubric (2 pts setup, 2 pts algebra, 2 pts answer)**

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Teacher: Mr. Hisham Mahmoud

Unit 7 Equilibrium – Comprehensive Review Answer Key

☑️ Use this key to check student work, award partial credit, and identify common mistakes!

Total: 100 points (30 MC + 40 Short Answer + 30 Calculations)