AP Chemistry Unit 7 Easier Practice Set

Equilibrium

Building Confidence with K expressions, ICE tables, and Le Chatelier

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APChemistryRescue.com

📋 Practice Set Information

- **Total Points:** 75 points
- **Suggested Time:** 70–100 minutes
- **Questions:** 10 (scaffolded from K expressions → ICE tables → Le Chatelier → calculations)
- **Topics:** Equilibrium expressions (Kc, Kp), homogeneous/heterogeneous equilibria, Q vs. K, ICE tables, Le Chatelier's principle, equilibrium calculations
- **Skills:** Writing K expressions, ICE table setup, predicting shifts, algebraic equilibrium calculations, Kp ↔ Kc conversions

🔬 ESSENTIAL EQUILIBRIUM FORMULAS & RULES

EQUILIBRIUM CONSTANT EXPRESSIONS:

```
For aA + bB ⇌ cC + dD
K_c = [C]^c[D]^d / [A]^a[B]^b (concentrations, M)
K_p = (P_C)^c(P_D)^d / (P_A)^a(P_B)^b (partial pressures, atm)
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IMPORTANT RULES:

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• Exclude pure solids (s) and pure liquids (l) from K
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- Include gases (g) and aqueous solutions (aq)
- K_p = K_c(RT)^Δn where Δn = (moles gas products) - (moles gas reactants)
- R = 0.0821 L·atm/(mol·K)

Q vs. K (REACTION QUOTIENT):

Q = same form as K, but uses CURRENT concentrations

- If Q < K → reaction shifts RIGHT (toward products)
- If Q > K → reaction shifts LEFT (toward reactants)
- If Q = K → system is at EQUILIBRIUM (no shift)

⚖ **Equilibrium Key Concepts:**

- **Dynamic equilibrium:**

  Forward and reverse rates are equal; concentrations constant (NOT zero)
- **K is temperature-dependent:**

  Changing T changes K; other stresses shift position but K stays same
- **Le Chatelier's Principle:**

  System responds to stress by shifting to relieve it
- **ICE tables:**

  Initial → Change (±x with stoichiometry) → Equilibrium

Question 1: Writing Equilibrium Expressions (Homogeneous) (6 points)

Write the equilibrium constant expressions (K_c) for the following **homogeneous** (all species in same phase) gas-phase reactions:

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

Work Space:

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

Work Space:

(c) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$

Work Space:

(d) For reaction (a), if the reaction is reversed ($2NH_3 \rightleftharpoons N_2 + 3H_2$), what is the relationship between the new K and the original K?

Work Space:

Question 2: Heterogeneous Equilibria (7 points)

Heterogeneous equilibria involve species in different phases. Pure solids (s) and pure liquids (l) are **excluded** from equilibrium expressions because their activities equal 1.

🔑 **Rule:**

Only include (g) and (aq) species in K expressions. Omit (s) and (l).

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

Write the K_c expression. Which species are included?

Work Space:

(b) $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$

Write the K_c expression.

Work Space:

(c) $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$

Write the K_c expression. (This is actually K_sp, solubility product, but follows the same rule.)

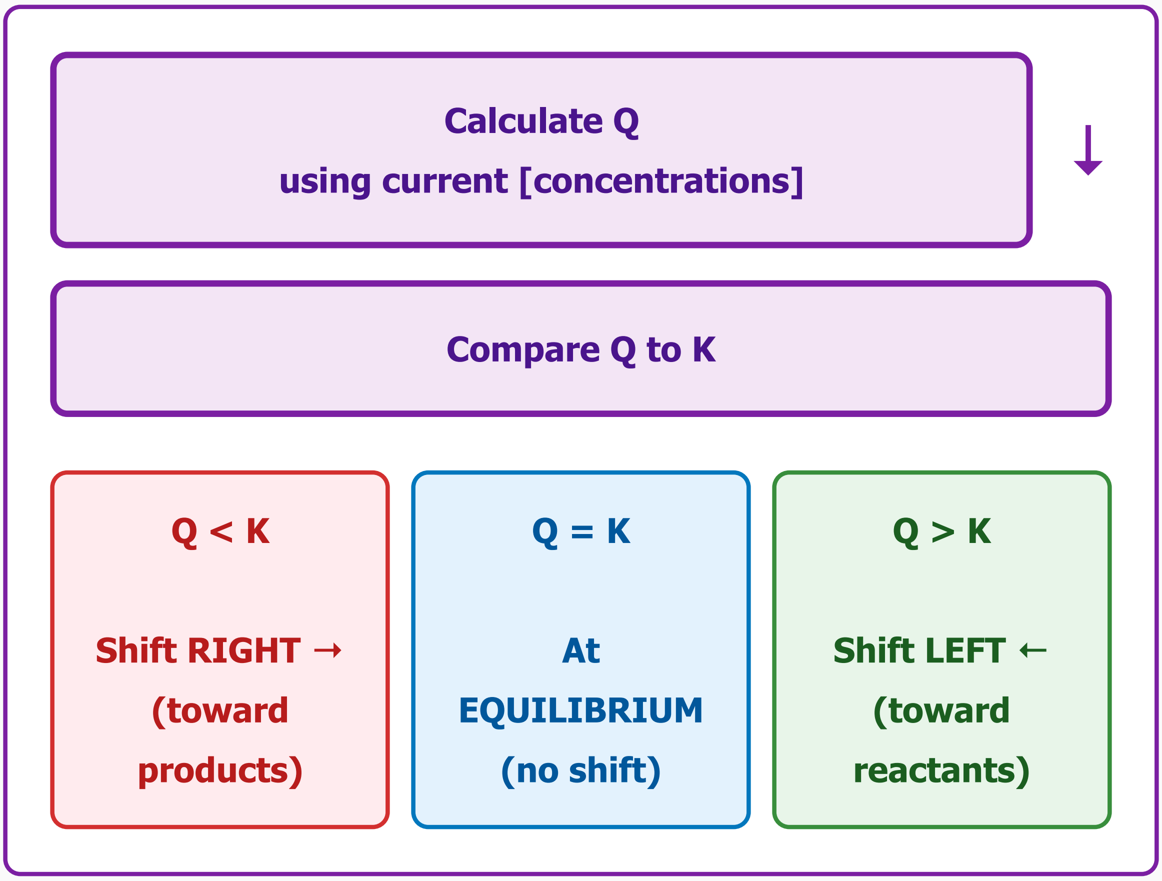
Work Space:

(d) Explain WHY pure solids and liquids are excluded from equilibrium expressions.

Work Space:

Question 3: Reaction Quotient (Q) vs. Equilibrium Constant (K) (8 points)

The **reaction quotient Q** has the same form as K but uses **current** (not equilibrium) concentrations. Comparing Q to K predicts the direction of shift:



Example: For the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, K_c = 0.500 at 400°C.

A reaction mixture contains $[N_2] = 0.10$ M, $[H_2] = 0.20$ M, and $[NH_3] = 0.040$ M.

(a) Write the expression for Q_c.

Work Space:

(b) Calculate Q_c using the given concentrations. Show all work.

Work Space:

(c) Compare Q_c to K_c. Is Q < K, Q = K, or Q > K?

Work Space:

(d) In which direction will the reaction shift to reach equilibrium? Will $[NH_3]$ increase or decrease as the system reaches equilibrium?

Work Space:

Question 4: ICE Table Setup (Basic) (8 points)

An **ICE table** (Initial, Change, Equilibrium) organizes equilibrium calculations.

Example: $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, K_c = 50.0 at 450°C

Initially, a container holds $[H_2]_0 = 1.00$ M, $[I_2]_0 = 1.00$ M, and $[HI]_0 = 0$ M.

ICE Table Rules:

- **Initial:**

  Write starting concentrations
- **Change:**

  Use variable x with stoichiometric coefficients (±1x, ±2x, etc.)
- **Equilibrium:**

  Add Initial + Change for each species

	$H_2(g)$	$I_2(g)$	$2HI(g)$
Initial (M)	1.00	1.00	0
Change (M)	$-x$	$-x$	$+2x$
Equilibrium (M)	$1.00 - x$	$1.00 - x$	$2x$

(a) Explain why the Change row shows $-x$ for H_2 and I_2, but $+2x$ for HI.

Work Space:

(b) Write the equilibrium expression K_c in terms of the equilibrium concentrations from the ICE table.

Work Space:

(c) Substitute the equilibrium expressions into $K_c = 50.0$ and set up the equation to solve for x. (You do NOT need to solve it yet.)

Work Space:

(d) Simplify the equation by taking the square root of both sides (since it's a perfect square setup). Then solve for x.

Hint:

$$\sqrt{50.0} = \sqrt{\frac{(2x)^2}{(1-x)^2}} = \frac{2x}{1-x}$$

Work Space:

Question 5: ICE Table with Small K Approximation (9 points)

For reactions with very small K (K << 1), we can use the **small K approximation**: $[A]_0 - x \approx [A]_0$ (if x is negligible).

`2NO₂(g) ⇌ 2NO(g) + O₂(g) K_c = 1.0 × 10⁻⁵ at 500 K`

Initially, $[NO_2]_0 = 0.500$ M, $[NO]_0 = 0$ M, $[O_2]_0 = 0$ M.

(a) Set up an ICE table for this reaction. Use variable x.

ICE Table (draw your own or use the template below):

	$2NO_2(g)$	2NO(g)	$O_2(g)$
Initial (M)			
Change (M)			
Equilibrium (M)			

(b) Write the K_c expression in terms of equilibrium concentrations from your ICE table.

Work Space:

(c) Apply the **small K approximation**: assume 0.500 − 2x ≈ 0.500 (because K is very small, x will be tiny). Solve for x.

Approximation:

$$K_c = \frac{[NO]^2[O_2]}{[NO_2]^2} \approx \frac{(2x)^2(x)}{(0.500)^2}$$

Work Space:

(d) Check the validity of the approximation by calculating $$\frac{2x}{0.500} \times 100\%$$. If this is less than 5%, the approximation is valid.

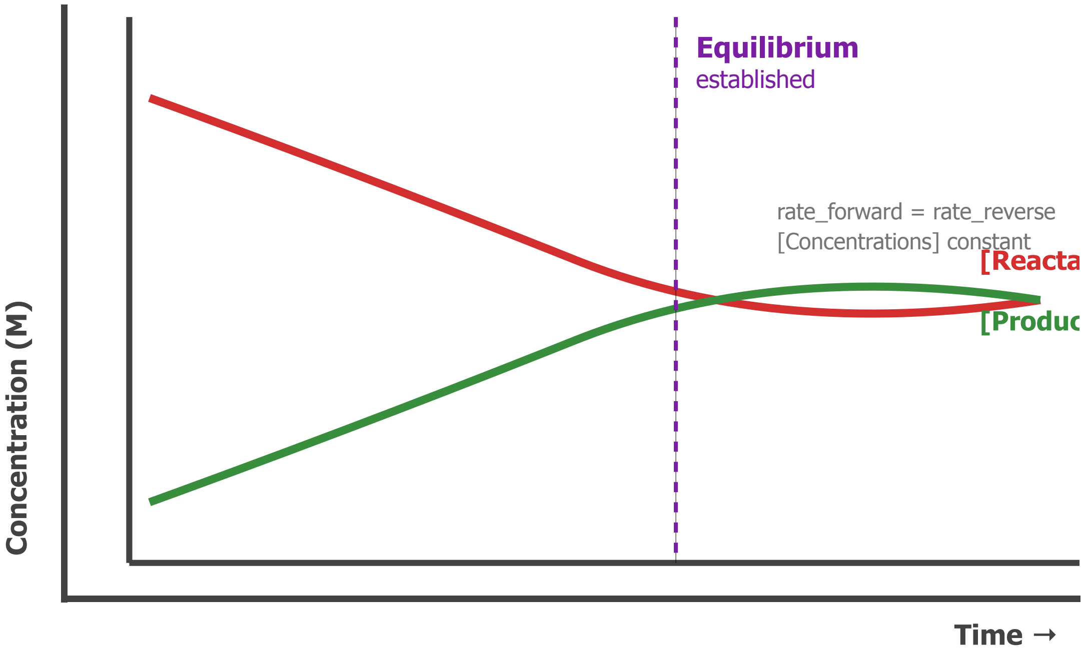
Work Space:

(e) Calculate the equilibrium concentration $[O_2]$_eq.

Work Space:

Question 6: Concentration vs. Time Graph (6 points)

When a reversible reaction reaches equilibrium, the concentrations of reactants and products become **constant** (but NOT zero).



(a) On the graph, why do both curves (reactants and products) **level off** after a certain time?

Work Space:

(b) Do the [reactants] and [products] need to be **equal** at equilibrium? Explain using the graph.

Work Space:

(c) What does it mean that equilibrium is **dynamic** (not static)?

Work Space:

Question 7: Le Chatelier's Principle (10 points)

Le Chatelier's Principle: If a stress is applied to a system at equilibrium, the system shifts in the direction that relieves the stress.

Consider the reaction: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92$ kJ (exothermic)

Stress Applied	Direction of Shift	Effect on $[NH_3]$
Add $N_2(g)$		
Remove $H_2(g)$		
Add $NH_3(g)$		
Increase pressure (decrease volume)		
Increase temperature		

(a) Complete the table above by predicting the direction of shift and effect on [NH_3] for each stress.

Work Space (you can fill in the table directly or write explanations here):

(b) Explain why increasing the **pressure** (by decreasing volume) shifts the equilibrium toward NH_3 in this reaction.

Hint:

Count moles of gas on each side of the equation.

Work Space:

(c) Explain why increasing the **temperature** shifts the equilibrium to the LEFT (toward reactants) for this exothermic reaction.

Work Space:

(d) Does adding a **catalyst** shift the equilibrium position or change the value of K? Explain.

Work Space:

Question 8: K_p and K_c Relationship (7 points)

For gas-phase reactions, the equilibrium constant can be expressed using concentrations (K_c) or partial pressures (K_p):

`K_p = K_c(RT)^Δn`

`Δn = (moles of gas products) - (moles of gas reactants)`

`R = 0.0821 L·atm/(mol·K)`

(a) For the reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$, calculate **Δn**.

Work Space:

(b) If K_c = 1.5 × 10² at 700 K, calculate K_p. Show all work with units.

Work Space:

(c) For the reaction $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, what is Δn? What is the relationship between K_p and K_c for this reaction?

Work Space:

Question 9: Calculating K from Equilibrium Concentrations (6 points)

At equilibrium at 500 K, a reaction mixture contains:

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

[CO] = 0.150 M, $[H_2]$ = 0.200 M, $[CH_3OH]$ = 0.0500 M

(a) Write the equilibrium expression K_c.

Work Space:

(b) Calculate K_c. Show all work with proper significant figures.

Work Space:

(c) Is this equilibrium favoring reactants or products? How do you know from the value of K?

Work Space:

Question 10: Conceptual Understanding (8 points)

(a) A student claims: "At equilibrium, the forward and reverse reactions stop." Is this correct? Explain what actually happens at equilibrium.

Work Space:

(b) Explain the difference between the **equilibrium position** shifting and the **equilibrium constant K** changing.

Work Space:

(c) Why can we not include pure solids and liquids in equilibrium expressions? (Hint: Consider their concentration or activity.)

Work Space:

(d) For an exothermic reaction, increasing temperature **decreases K**. Explain this using Le Chatelier's principle and the relationship between heat and equilibrium.

Work Space:

END OF PRACTICE SET

Total: 75 points | Answer key begins on next page

📘 COMPLETE ANSWER KEY & SCORING GUIDE

Question 1: Writing Equilibrium Expressions (6 points)

(a) $N_2 + 3H_2 \rightleftharpoons 2NH_3$ (1.5 points):

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

(1 pt for correct form; 0.5 pt for correct exponents)

(b) $2SO_2 + O_2 \rightleftharpoons 2SO_3$ (1.5 points):

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

(c) $PCl_5 \rightleftharpoons PCl_3 + Cl_2$ (1.5 points):

$$K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]}$$

(d) Relationship when reversed (1.5 points):

When a reaction is reversed, the new K is the **reciprocal** of the original K. (1 pt)

$$K_{reverse} = \frac{1}{K_{forward}}$$

For reaction (a), if K_forward = x, then K_reverse = 1/x. (0.5 pt)

Scoring: (a) 1.5 pts; (b) 1.5 pts; (c) 1.5 pts; (d) 1.5 pts

Question 2: Heterogeneous Equilibria (7 points)

(a) $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ (2 points):

$$K_c = [CO_2]$$

Only $CO_2(g)$ is included. (1 pt) $CaCO_3(s)$ and $CaO(s)$ are pure solids, so they are excluded. (1 pt)

(b) $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$ (2 points):

$$K_c = \frac{[CO][H_2]}{[H_2O]}$$

C(s) is excluded because it's a pure solid. (2 pts: 1 for expression, 1 for excluding C)

(c) $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$ (1.5 points):

$$K_c = K_{sp} = [Ag^+][Cl^-]$$

AgCl(s) is excluded. (1.5 pts)

(d) Why exclude solids and liquids (1.5 points):

The **concentration** of a pure solid or pure liquid is essentially constant—it doesn't change during the reaction because its density is fixed. (1 pt) In thermodynamic terms, the **activity** of a pure solid or liquid is defined as 1, so including it in the K expression would just multiply by 1 (no effect). Therefore, we exclude them for simplicity and focus only on species whose concentrations can vary (gases and aqueous solutions). (0.5 pt)

Scoring: (a) 2 pts; (b) 2 pts; (c) 1.5 pts; (d) 1.5 pts

Question 3: Q vs. K (8 points)

(a) Expression for Q_c (2 points):

$$Q_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

(Same form as K_c, but uses current concentrations.) (2 pts)

(b) Calculate Q_c (3 points):

$$Q_c = \frac{(0.040)^2}{(0.10)(0.20)^3} = \frac{0.0016}{(0.10)(0.008)} = \frac{0.0016}{0.0008} = 2.0$$

(1 pt for substitution; 1 pt for calculation; 1 pt for answer)

(c) Compare Q to K (1 point):

Q_c = 2.0, K_c = 0.500

Q > K (1 pt)

(d) Direction of shift (2 points):

Because Q > K, the reaction will shift **LEFT** (toward reactants) to reach equilibrium. (1 pt) This means [NH_3] will **decrease** (consumed to form more N_2 and H_2). (1 pt)

Scoring: (a) 2 pts; (b) 3 pts; (c) 1 pt; (d) 2 pts

Question 4: ICE Table Setup (8 points)

(a) Why −x and +2x? (2 points):

From the balanced equation H_2 + I_2 → 2HI, for every 1 mole of H_2 that reacts, 1 mole of I_2 also reacts, and 2 moles of HI are formed. (1 pt) The stoichiometric coefficients dictate the Change row: H_2 and I_2 decrease by x each (−x), while HI increases by 2x (+2x) to maintain the 1:1:2 ratio. (1 pt)

(b) K_c expression (1.5 points):

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(2x)^2}{(1.00-x)(1.00-x)}$$

(1.5 pts for correct substitution)

(c) Set up equation (1.5 points):

$$50.0 = \frac{(2x)^2}{(1.00-x)^2} = \frac{4x^2}{(1.00-x)^2}$$

(1.5 pts)

(d) Solve for x (3 points):

Take the square root of both sides:

$$\sqrt{50.0} = \frac{2x}{1.00-x}$$

$$7.07 = \frac{2x}{1.00-x}$$

$$7.07(1.00-x) = 2x$$

$$7.07 - 7.07x = 2x$$

$$7.07 = 9.07x$$

$$x = \frac{7.07}{9.07} = 0.780$$

(1 pt for taking square root; 1 pt for algebra; 1 pt for final answer)

Equilibrium concentrations: $[H_2] = [I_2] = 1.00 - 0.780 = 0.220$ M; [HI] = 2(0.780) = 1.56 M

Scoring: (a) 2 pts; (b) 1.5 pts; (c) 1.5 pts; (d) 3 pts

Question 5: ICE Table with Small K Approximation (9 points)

(a) ICE table (2 points):

	$2NO_2(g)$	$2NO(g)$	$O_2(g)$
Initial (M)	**0.500**	**0**	**0**
Change (M)	**−2x**	**+2x**	**+x**
Equilibrium (M)	**0.500 − 2x**	**2x**	**x**

(1 pt for correct Initial; 1 pt for correct Change with stoichiometry)

(b) K_c expression (1.5 points):

$$K_c = \frac{[NO]^2[O_2]}{[NO_2]^2} = \frac{(2x)^2(x)}{(0.500-2x)^2}$$

(1.5 pts)

(c) Solve using approximation (3 points):

Approximation: 0.500 − 2x ≈ 0.500

$$1.0 \times 10^{-5} = \frac{(2x)^2(x)}{(0.500)^2} = \frac{4x^3}{0.250}$$

$$1.0 \times 10^{-5} = 16x^3$$

$$x^3 = \frac{1.0 \times 10^{-5}}{16} = 6.25 \times 10^{-7}$$

$$x = \sqrt[3]{6.25 \times 10^{-7}} = 8.55 \times 10^{-3}$$

(1 pt for setting up approximation; 1 pt for algebra; 1 pt for solving x)

(d) Check validity (1.5 points):

$$\frac{2x}{0.500} \times 100\% = \frac{2(8.55 \times 10^{-3})}{0.500} \times 100\% = \frac{0.0171}{0.500} \times 100\% = 3.42\%$$

Since 3.42% < 5%, the approximation is **valid**. (1 pt for calculation; 0.5 pt for conclusion)

(e) Equilibrium [O_2] (1 point):

$$[O_2]_{eq} = x = 8.55 \times 10^{-3} \text{ M}$$

(1 pt)

Scoring: (a) 2 pts; (b) 1.5 pts; (c) 3 pts; (d) 1.5 pts; (e) 1 pt

Question 6: Concentration vs. Time Graph (6 points)

(a) Why curves level off (2 points):

The curves level off because the system has reached **equilibrium**. (1 pt) At equilibrium, the forward reaction rate equals the reverse reaction rate, so the concentrations of reactants and products remain constant (no net change). (1 pt)

(b) Must [reactants] = [products]? (2 points):

No, [reactants] and [products] do NOT need to be equal at equilibrium. (1 pt) The graph shows [products] and [reactants] reaching different constant values. Equilibrium means the **rates** are equal, not the concentrations. The ratio of concentrations is determined by K, which can be large (product-favored) or small (reactant-favored). (1 pt)

(c) Dynamic equilibrium (2 points):

Dynamic equilibrium means the forward and reverse reactions **continue to occur** at the molecular level, but at equal rates. (1 pt) Molecules are constantly reacting in both directions, but because the rates are balanced, the macroscopic concentrations don't change—it's not a "static" or "frozen" state. (1 pt)

Scoring: (a) 2 pts; (b) 2 pts; (c) 2 pts

Question 7: Le Chatelier's Principle (10 points)

(a) Completed table (5 points, 1 pt per row):

Stress Applied	Direction of Shift	Effect on [NH_3]
Add N_2(g)	**Shift RIGHT (→)**	Increase
Remove H_2(g)	**Shift LEFT (←)**	Decrease
Add NH_3(g)	**Shift LEFT (←)**	Decrease (initially increases, then system consumes some)
Increase pressure (decrease volume)	**Shift RIGHT (→)**	Increase
Increase temperature	**Shift LEFT (←)**	Decrease

(b) Why pressure increase shifts right (2 points):

The left side has 1 mol N_2 + 3 mol H_2 = **4 moles of gas**. The right side has **2 moles of gas** (NH_3). (1 pt) When pressure increases (volume decreases), the system shifts toward the side with **fewer moles of gas** to reduce pressure—in this case, toward NH_3 (right). (1 pt)

(c) Why temperature increase shifts left (2 points):

The reaction is exothermic ($\Delta H = -92$ kJ), meaning heat is a "product": $N_2 + 3H_2 \rightleftharpoons 2NH_3$ + heat. (1 pt) Increasing temperature adds heat (a stress). By Le Chatelier, the system shifts to consume the added heat, which means shifting LEFT (toward reactants), treating heat as a product to be removed. This also means K **decreases** with increasing T for exothermic reactions. (1 pt)

(d) Catalyst effect (1 point):

Adding a catalyst **does NOT shift the equilibrium position** or change K. (0.5 pt) A catalyst speeds up both forward and reverse reactions equally, so the system reaches equilibrium **faster**, but the final equilibrium concentrations remain the same. (0.5 pt)

Scoring: (a) 5 pts (1 per row); (b) 2 pts; (c) 2 pts; (d) 1 pt

Question 8: K_p and K_c Relationship (7 points)

(a) Calculate Δn (2 points):

$2SO_2 + O_2 \rightleftharpoons 2SO_3$

Reactants: 2 mol SO_2 + 1 mol O_2 = 3 moles gas

Products: 2 mol SO_3 = 2 moles gas

$$\Delta n = 2 - 3 = -1$$

(1 pt for calculation; 1 pt for correct answer)

(b) Calculate K_p (3 points):

$$K_p = K_c(RT)^{\Delta n} = (1.5 \times 10^2)(0.0821 \times 700)^{-1}$$

$$= (1.5 \times 10^2) \times \frac{1}{(0.0821)(700)} = \frac{150}{57.47} = 2.61$$

(1 pt for formula; 1 pt for substitution; 1 pt for answer)

(c) Δn for $H_2 + I_2 \rightleftharpoons 2HI$ (2 points):

Reactants: 1 mol H_2 + 1 mol I_2 = 2 moles gas

Products: 2 mol HI = 2 moles gas

$$\Delta n = 2 - 2 = 0$$

When Δn = 0, **K_p = K_c** (because $(RT)^0 = 1$). (1 pt for Δn; 1 pt for relationship)

Scoring: (a) 2 pts; (b) 3 pts; (c) 2 pts

Question 9: Calculating K from Equilibrium Concentrations (6 points)

(a) K_c expression (1.5 points):

$$K_c = \frac{[CH_3OH]}{[CO][H_2]^2}$$

(1.5 pts)

(b) Calculate K_c (3 points):

$$K_c = \frac{0.0500}{(0.150)(0.200)^2} = \frac{0.0500}{(0.150)(0.0400)} = \frac{0.0500}{0.00600} = 8.33$$

(1 pt for substitution; 1 pt for calculation; 1 pt for sig figs)

(c) Reactants or products favored? (1.5 points):

K_c = 8.33 is **greater than 1**, so the equilibrium **favors products** (numerator > denominator). (1 pt) When K >> 1, products predominate at equilibrium; when K << 1, reactants predominate. (0.5 pt)

Scoring: (a) 1.5 pts; (b) 3 pts; (c) 1.5 pts

Question 10: Conceptual Understanding (8 points)

(a) Do reactions stop at equilibrium? (2 points):

Incorrect. At equilibrium, the forward and reverse reactions **continue to occur**, but at equal rates. (1 pt) This is called **dynamic equilibrium**—molecules are constantly reacting in both directions, but the concentrations remain constant because the rates balance. The reactions never "stop." (1 pt)

(b) Equilibrium position vs. K changing (2 points):

The **equilibrium position** refers to the specific concentrations (or ratio) of reactants and products at equilibrium, which can shift in response to stresses like concentration changes, pressure, etc. (1 pt) The **equilibrium**

constant K is a numerical value that depends **only on temperature**. K changes only if temperature changes. Other stresses (adding reactants, changing pressure) shift the position but do NOT change K. (1 pt)

(c) Why exclude solids and liquids? (2 points):

The concentration (or activity) of a pure solid or pure liquid is **constant**—it doesn't change during the reaction because the density and amount per unit volume are fixed. (1 pt) Mathematically, their activity is defined as 1, so including them in the K expression would just multiply by 1 (no effect). We exclude them for simplicity, focusing only on species whose concentrations vary: gases (g) and aqueous solutions (aq). (1 pt)

(d) Why increasing T decreases K for exothermic reactions (2 points):

For an exothermic reaction, heat is a "product": Reactants $\rightleftharpoons$ Products + heat. (1 pt) Increasing temperature adds heat (a stress). By Le Chatelier's principle, the system shifts to consume the added heat by favoring the reverse reaction (toward reactants), which decreases the [products]/[reactants] ratio—thus K decreases. Conversely, for endothermic reactions, increasing T increases K because the system shifts toward products to absorb the added heat. (1 pt)

Scoring: (a) 2 pts; (b) 2 pts; (c) 2 pts; (d) 2 pts