This print-out should have 27 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001 4.0 points

A 10.0 L vessel contains 0.0015 mole CO_2 and 0.10 mole CO. If a small amount of carbon is added to this vessel and the temperature is raised to 1000°C

$$CO_2(g) + C(s) \rightleftharpoons 2\,CO(g)\,,$$

will more CO form? The value of K_c for this reaction is 1.17 at 1000°C. Assume that the volume of the gas in the vessel is 10.0 L.

1. Yes, the rate of the forward reaction will increase to produce more CO. **correct**

2. Unable to determine this from the data provided.

3. No, the rate of the reverse reaction will increase to produce more CO_2.

Explanation:

$[CO] = \dfrac{0.1 \text{ mol}}{10 \text{ L}}$ $\qquad [CO_2] = \dfrac{0.0015 \text{ mol}}{10 \text{ L}}$

Carbon, being a solid, has no effect on equilibrium.

$$[Q] = \frac{[CO]^2}{[CO_2]} = \frac{\left(\dfrac{0.1}{10.0}\text{ M}\right)^2}{\left(\dfrac{0.0015}{10.0}\text{ M}\right)}$$
$$= 0.666667 < K_c = 1.17$$

Therefore equilibrium will shift to the right.

002 4.0 points

The expression for K_c for the reaction at equilibrium is

$4\text{ }NH_3(g) + 5\text{ }O_2(g) \rightleftharpoons 4\text{ }NO(g) + 6\text{ }H_2O(g)$

1. $\dfrac{[NH_3]^4\,[O_2]^5}{[NO]^4\,[H_2O]^6}$

2. $[NO]^4\,[H_2O\,]^6$

3. $[NH_3]^4\,[O_2]^5$

4. $\dfrac{[NO]^4\,[H_2O]^6}{[NH_3]^4\,[O_2]^5}$ **correct**

Explanation:

The equation must be written with the appropriate formula and correctly balanced. K_c is the equilibrium constant for species in solution and equals the mathematical product of the concentrations of the chemical products, divided by the mathematical product of the concentrations of the chemical reactants. In this mathematical expression, each concentration is raised to the power of its coefficient in the balanced equation. For K_c the molar concentrations are used for the activities of the components.

003 4.0 points

Consider the reaction

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

where $\Delta H_{rxn} = -198$ kJ. The amount of $SO_2(g)$ at equilibrium increases when

1. the pressure is increased.

2. the volume is increased. **correct**

3. the temperature is decreased.

4. more oxygen is added.

5. SO_3 is removed.

Explanation:

$$2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_2(g) + \text{heat}$$

According to Le Chatelier's principle, the amount of reactant $SO_2(g)$ is increased when the equilibrium shifts to the left. This will happen when another reactant (O_2) is removed. For an exothermic reaction decreasing temperature removes heat and sends equilibrium to the right. Increasing pressure sends the equilibrium in the direction that has fewer numbers of moles of gas. Increasing the volume is equivalent to decreasing gas pressure.

004 4.0 points

The reaction

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

has an equilibrium constant of 4.0×10^8 at 25°C. What will eventually happen if 44.0 moles of NH_3, 0.452 moles of N_2, and 0.108 moles of H_2 are put in a 10.0 liter container at 25°C?

1. More NH_3 will be formed. **correct**

2. More N_2 and H_2 will be formed.

3. Nothing; the system is at equilibrium.

Explanation:

$K = 4.0 \times 10^8$ $\qquad [NH_3] = \dfrac{44.0\text{ mol}}{10\text{ L}}$

$[N_2] = \dfrac{0.452\text{ mol}}{10\text{ L}}$ $\qquad [H_2] = \dfrac{0.108\text{ mol}}{10\text{ L}}$

$$\begin{aligned} Q &= \frac{[NH_3]^2}{[N_2]\,[H_2]^3} \\ &= \frac{(4.40\text{ M})^2}{(0.0452\text{ M})\,(0.0108\text{ M})^3} \\ &= 3.4 \times 10^8 \end{aligned}$$

Since $Q < K$ equilibrium will shift to the right, forming more NH_3.

005 4.0 points

The reaction

$$A + B \rightleftharpoons C + 2\,D$$

has an equilibrium constant of 3.7×10^{-3}. Consider a reaction mixture with

$[A] = 2.0 \times 10^{-2}$ M $\qquad [C] = 2.4 \times 10^{-6}$ M

$[B] = 1.7 \times 10^{-4}$ M $\qquad [D] = 3.5 \times 10^{-3}$ M

Which of the following statements is definitely true?

1. The forward reaction can occur to a greater extent than the reverse reaction until equilibrium is established. **correct**

2. No conclusions about the system can be made without additional information.

3. The reverse reaction can occur to a greater extent than the forward reaction until equilibrium is established.

4. Heat will be evolved.

5. The system is at equilibrium.

Explanation:

$$\begin{aligned} Q &= \frac{[C]\,[D]^2}{[A]\,[B]} = \frac{(2.4 \times 10^{-6}\text{ M})\,(0.0035\text{ M})^2}{(0.02\text{ M})\,(0.00017\text{ M})} \\ &= 8.64706 \times 10^{-6} \end{aligned}$$

Since $Q < K$ the foward reaction is favored.

006 4.0 points

For an exothermic reaction, what would happen to the numerical value of K_c, if we increase the temperature at constant pressure?

1. K_c would either increase or decrease, depending on the number of moles of gas involved.

2. K_c would not change.

3. K_c would decrease. **correct**

4. K_c would either increase or decrease, depending on the concentrations.

5. K_c would increase.

Explanation:

An exothermic reaction gives off heat as products are formed, so at a higher temperature the reverse endothermic reaction is favoured, decreasing K_c, which reflects the ratio of the products to reactants.

007 4.0 points

Suppose the reaction mixture

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

is at equilibrium at a given temperature and pressure. The pressure is then increased at constant temperature by compressing the reaction mixture, and the mixture is then allowed to reestablish equilibrium. At the new equilibrium,

1. there is the same amount of ammonia present as there was originally.

2. the nitrogen is used up completely.

3. there is more ammonia present than there was originally. **correct**

4. there is less ammonia present than there was originally.

Explanation:
LeChatelier's Principle states that if a change occurs in a system at equilibruim, the system responds to relieve the stress and reach a new equilibrium. Here, the number of moles of gaseous reactants is greater than the number of moles of products. Increasing the pressure of the above system will result in the reaction proceeding to reduce that pressure increase. The system will shift to the right (the side that has fewer moles of gas), so the pressure will be reduced; thus more ammonia will be produced.

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008 4.0 points

What happens to the concentration of NO(g) when the total pressure on the equilibrium reaction

$$3\,NO_2(g) + H_2O(\ell) \rightleftharpoons 2\,HNO_3(aq) + NO(g)$$

is increased (by compression)?

1. decreases

2. increases **correct**

3. remains the same

4. Unable to determine

Explanation:
Increasing the total pressure on the system by decreasing its volume will shift the equilibrium toward the side of the reaction with fewer numbers of moles of gaseous components. If the total number of moles of gas is the same on the product and reactant sides of the balanced chemical equation, then changing the pressure will have little or no effect on the equilibrium distribution of species present. The amount and concentration of NO will increase.

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009 4.0 points
Consider the system

$$2\,N_2O_5(g) \rightleftharpoons 2\,N_2O_4(g) + O_2(g) + \text{heat}$$

at equilibrium at 25°C. If the temperature were raised would the equilibrium be shifted to produce more N_2O_5 or more N_2O_4?

1. There would be no effect.

2. more N_2O_5 **correct**

3. more N_2O_4

Explanation:
This is an exothermic reaction and so increasing temperature provides more heat to the system, shifting equilibrium to the left and producing more N_2O_5.

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010 4.0 points
The system

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

is at equilibrium at a fixed temperature with a partial pressure of H_2 of 0.200 atm, a partial pressure of I_2 of 0.200 atm, and a partial pressure of HI of 0.100 atm. An additional 0.34 atm pressure of HI is admitted to the container, and it is allowed to come to equilibrium again. What is the new partial pressure of HI?

1. 0.142
2. 0.152
3. 0.138
4. 0.132
5. 0.162
6. 0.164
7. 0.168

8. 0.146
9. 0.15
10. 0.136

Correct answer: 0.168 atm.

Explanation:
$P_{I_2} = P_{H_2} = 0.2$ atm $\qquad P_{HI} = 0.1$ atm

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

$$\begin{aligned} K_p &= \frac{(P_{HI})^2}{P_{H_2} \cdot P_{I_2}} \\ &= \frac{(0.1 \text{ atm})^2}{(0.2 \text{ atm})\,(0.2 \text{ atm})} = 0.25 \\ \text{new } P_{HI} &= (0.1 + 0.34) \text{ atm} = 0.44 \text{ atm} \end{aligned}$$

Adding the products shifts the equilibrium to the left.

	$H_2(g)$	+ $I_2(g)$	$\rightleftharpoons$ $2\,HI(g)$
ini, atm	0.200	0.200	0.44
Δ, atm	$+x$	$+x$	$-2x$
eq, atm	$0.200 + x$	$0.200 + x$	$0.44 - 2x$

$$\begin{aligned} \frac{(0.44 - 2\,x)^2}{(0.2 + x)^2} &= 0.25 \\ \frac{0.44 - 2\,x}{0.2 + x} &= \sqrt{0.25} \\ 0.44 - 2\,x &= (0.5)(0.2) + 0.5\,x \\ 2.5\,x &= 0.34 \\ x &= 0.136 \end{aligned}$$

$$P_{HI} = 0.44 \text{ atm} - (2)\,(0.136 \text{ atm}) = 0.168 \text{ atm}$$

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011 4.0 points

Consider the following reactions at 25°C:

reaction	K_c
$2\,NO(g) \rightleftharpoons N_2(g) + O_2(g)$	1×10^{30}
$2\,H_2O(g) \rightleftharpoons 2\,H_2(g) + O_2(g)$	5×10^{-82}
$2\,CO(g) + O_2(g) \rightleftharpoons 2\,CO_2(g)$	3×10^{91}

Which compound is most likely to dissociate and give $O_2(g)$ at 25°C?

1. CO

2. CO_2

3. NO **correct**

4. H_2O

Explanation:
Only two are dissociation reactions: dissociation of NO and dissociation of H_2O. K_c is greater for dissociation of NO.

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012 4.0 points

Suppose the reaction

$$A \rightleftharpoons B$$

has an equilibrium constant of 1.0 and the initial concentrations of A and B are 0.5 M and 0.0 M, respectively. Which of the following is the correct value for the equilibrium concentration of A?

1. None of these is correct.

2. 0.250 M **correct**

3. 1.50 M

4. 0.500 M

5. 1.00 M

Explanation:
$K = 1.0 \qquad [A]_{ini} = 0.5$ M
$[B]_{ini} = 0$ M

	A	$\rightleftharpoons$ B
ini, M	0.5	0.0
Δ, M	$-x$	x
eq, M	$0.5 - x$	x

$$\begin{aligned} K &= \frac{[B]}{[A]} = 1.0 \\ \frac{x}{0.5 - x} &= 1.0 \\ x &= 0.25 \text{ M} \end{aligned}$$

$$[A] = 0.5 - x = 0.25 \text{ M}$$

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013 4.0 points

The system

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

is at equilibrium at some temperature. At equilibrium a 4.00 L vessel contains 1.00 mole CO_2, 1.00 mole H_2, 2.40 moles H_2O, and 2.40 moles CO. How many moles of CO_2 must be added to the system to bring the equilibrium CO concentration to 0.661 mol/L?

1. 1.188
2. 1.069
3. 1.694
4. 8.112
5. 0.732
6. 3.672
7. 0.849
8. 6.576
9. 2.121
10. 1.747

Correct answer: 0.849 moles.

Explanation:

$V = 4.0$ L $\quad n_{CO_2} = 1.0$ mol

$n_{H_2} = 1.0$ mol $\quad n_{H_2O} = 2.40$ mol

$n_{CO} = 2.40$ mol

At equilibrium, the concentrations are

$$[CO_2] = [H_2] = \frac{1\text{ mol}}{4\text{ L}} = 0.25\text{ M}$$

$$[H_2O] = [CO] = \frac{2.4\text{ mol}}{4\text{ L}} = 0.6\text{ M}$$

First calculate the value of K:

$$K = \frac{[CO]\,[H_2O]}{[CO_2]\,[H_2]} = \frac{(0.6\text{ M})\,(0.6\text{ M})}{(0.25\text{ M})\,(0.25\text{ M})} = 5.76$$

Set up your equilibrium table knowing that the initial CO_2 concentration is now higher than 0.25 M. The reaction will go to the right and you can get the following...

CO_2	+	H_2	$\rightleftharpoons$	H_2O	+	CO
?		0.25		0.6		0.6
$-x$		$-x$		$+x$		$+x$
$?-x$		$0.25-x$		$0.6+x$		$0.6+x$

The final concentration of CO is 0.661, so

$$0.6 + x = 0.661$$
$$x = 0.661 - 0.6 = 0.061$$

and $[H_2] = 0.25 - x = 0.189$ M.

We also now know all of the final concentrations except CO_2, which is calculated from the equilibrium expression

$$5.76 = \frac{(0.661\text{ M})\,(0.661\text{ M})}{[CO_2]\,(0.189\text{ M})}$$

$$[CO_2] = \frac{(0.661\text{ M})\,(0.661\text{ M})}{(5.76\text{ M})\,(0.189\text{ M})}$$

so that $[CO_2]_{\text{fin}} = 0.4013$ mol/L, which corresponds to $0.4013 + 0.061 = 0.4623$ for the initial concentration of CO_2 BEFORE the equilibrium occurred.

That value corresponds to 1.849 moles of CO_2. Now substract the 1.00 mole of CO_2 that was already there and you get 0.849 moles of CO_2 that had to be added.

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014 4.0 points

Given the reaction

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

at equilibrium, if the pressure is doubled (think of the volume of the container halving), in which direction will the reaction shift?

1. left **correct**

2. no change

3. right

Explanation:

Increasing pressure shifts the equilibrium in the direction that produces fewer moles of gas.

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015 4.0 points

Consider the reaction

$$2\,HgO(s) \rightleftharpoons 2\,Hg(\ell) + O_2(g)\,.$$

What is the form of the equilibrium constant K for the reaction?

1. $K = \dfrac{[O_2]}{[HgO]^2}$

2. $K = [O_2]$ **correct**

3. $K = \dfrac{[Hg]^2\,[O_2]}{[HgO]^2}$

4. None of the other answers is correct.

5. $K = [\text{Hg}]^2\,[\text{O}_2]$

Explanation:

Solids and liquids are not included in the K expression.

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016 4.0 points

At 990°C, $K_c = 2.05$ for the reaction:

$$\text{H}_2(\text{g}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{H}_2\text{O}(\text{g}) + \text{CO}(\text{g})$$

How many moles of $H_2O(g)$ are present in an equilibrium mixture resulting from the addition of 1.38 mole of H_2, 1.89 moles of CO_2, 0.881 moles of H_2O, and 1.01 mole of CO to a 5.00 liter container at 990°C?

1. 1.44 mol **correct**

2. 1.50 mol

3. 1.56 mol

4. 1.41 mol

5. 1.47 mol

Explanation:

$K_c = 2.05$

$$Q = \frac{[\text{H}_2\text{O}][\text{CO}]}{[\text{H}_2][\text{CO}_2]} = \frac{\left(\frac{0.881\text{ mol}}{5.00\text{ L}}\right)\left(\frac{1.01\text{ mol}}{5.00\text{ L}}\right)}{\left(\frac{1.38\text{ mol}}{5.00\text{ L}}\right)\left(\frac{1.89\text{ mol}}{5.00\text{ L}}\right)} = 0.341$$

$0.341 < 2.05$, so the equilibrium moves to the right.

$H_2(g)$ +	$CO_2(g)$ ⇌	$H_2O(g)$ +	$CO(g)$
0.276 M	0.378 M	0.176 M	0.202 M
$-x$	$-x$	$+x$	$+x$
$0.276 - x$	$0.378 - x$	$0.176 + x$	$0.202 + x$

$$\frac{(0.176 + x)(0.202 + x)}{(0.276 - x)(0.378 - x)} = 2.05$$

Which reduces to: $1.05x^2 - 1.7189x + 0.17828 = 0$

$$x = \frac{1.7189 \pm \sqrt{(1.7189)^2 - 4(1.05)(0.17828)}}{2(1.05)} = 0.1113\text{ M}$$

$$[\text{H}_2\text{O}]_f = 0.1762\text{ M} + 0.1113\text{ M} = 0.2875\text{ M}$$

$$\text{mol H}_2\text{O} = 5.00\text{ L} \times 0.2875\,\frac{\text{mol}}{\text{L}} = 1.44\,\text{mol H}_2\text{O}$$

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017 4.0 points

For a certain reaction, $K = 44.6$ at 300 K and the reaction is endothermic by 7.3 kJ/mol. What is K at 500 K?

1. 98.2928
2. 107.698
3. 313.761
4. 407.498
5. 517.733
6. 354.101
7. 476.7
8. 179.632
9. 122.704
10. 143.804

Correct answer: 143.804.

Explanation:

$T_1 = 300$ K $\quad K_{300} = 44.6$
$T_2 = 500$ K $\quad \Delta H = 7.3$ kJ/mol
$K_{500} = ?$

$$\ln\frac{K_2}{K_1} = \ln K_2 - \ln K_1 = \frac{\Delta H}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\ln K_2 = \ln K_1 + \frac{\Delta H}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\ln K_{500} = \ln 44.6 + \frac{7300\text{ J/mol}}{8.314\text{ J/mol K}} \times \left(\frac{1}{300\text{ K}} - \frac{1}{500\text{ K}}\right) = 4.96845$$

$$K = e^{4.96845} = 143.804$$

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018 4.0 points

Given $K_p = 4.6 \times 10^{-14}$ and $\Delta H^0 = 115$ kJ/mol for the reaction

$$2\,Cl_2(g) + 2\,H_2O(g) \rightleftharpoons 4\,HCl(g) + O_2(g)$$

at 25°C, what is K_p at 400°C?

1. 1.4×10^{-5}

2. 7.7×10^{-3} **correct**

3. 3.9×10^{-4}

4. 7.9×10^{-2}

5. 1.3×10^{2}

Explanation:
T_1= 298.15 K T_2= 673.15 K

Use the van't Hoff equation:

$$\ln\left(\frac{K_2}{K_1}\right) = \frac{\Delta H^0_{\rm rxn}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$
$$\ln K_2 - \ln K_1 = \frac{\Delta H^0_{\rm rxn}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\begin{aligned}\ln K_2 &= \frac{\Delta H^0_{\rm rxn}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right) + \ln K_1\\ &= \left(\frac{115000\ \text{J/mol}}{8.314\ \text{J/mol}\cdot\text{K}}\right)\\ &\quad\times\left(\frac{1}{298.15\ \text{K}} - \frac{1}{673.15\ \text{K}}\right)\\ &\quad + \ln\left(4.6\times10^{-14}\right)\\ &= -4.86538\\ K_2 &= e^{-4.86538} = 0.0077089\end{aligned}$$

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019 4.0 points

Consider the reaction

$$\text{Ni(CO)}_4(g) \rightarrow \text{Ni(s)} + 4\,\text{CO}(g)\,.$$

If the initial concentration of $Ni(CO)_4(g)$ is 1.0 M, and x is the equilibrium concentration of CO(g), what is the correct equilibrium relation?

1. $K_c = \dfrac{x^5}{1.0 - \dfrac{x}{4}}$

2. $K_c = \dfrac{x^4}{1.0 - \dfrac{x}{4}}$ **correct**

3. $K_c = \dfrac{x}{1.0 - \dfrac{x}{4}}$

4. $K_c = \dfrac{4\,x}{1.0 - 4x}$

5. $K_c = \dfrac{x^4}{1.0 - 4x}$

Explanation:

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020 4.0 points

An equilibrium in which processes occur continuously, with NO NET change, is called

1. static equilibrium.

2. heterogeneous equilibrium.

3. homogeneous equilibrium.

4. dynamic equilibrium. **correct**

Explanation:
For a system at dynamic equilibrium, although the concentrations of the components do not change, the processes continue to occur in the foward and reverse directions at the same rate.

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021 4.0 points

At 1000 K the equilibrium pressure of the three gases in one mixture

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

were found to be 0.562 atm SO_2, 0.101 atm O_2, and 0.332 atm SO_3. Calculate the value of K_p for the reaction as written.

1. 0.289

2. 0.171

3. 5.83

4. 3.46 **correct**

5. 2.64

Explanation:
$P_{SO_3} = 0.332$ atm $\quad P_{SO_2} = 0.562$ atm
$P_{O_2} = 0.101$ atm

$$K_p = \frac{P^2_{SO_3}}{P^2_{SO_2} \cdot P_{O_2}} = \frac{(0.332)^2}{(0.562)^2(0.101)} = 3.46$$

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022 4.0 points

Calculate the equilibrium constant at 25°C for a reaction for which $\Delta G^0 = -3.35$ kcal/mol.

1. -285.64

2. 142.82

3. 571.281

4. 285.64 **correct**

5. 2856.4

Explanation:
$T = 25°\text{C} + 273 = 298$ K
$\Delta G^0 = -3350$ cal/mol
At equilibrium

$$\begin{aligned} \Delta G^0 &= -RT \ln K \\ -3350 &= (-1.987 \text{ cal/mol} \cdot \text{K}) \\ &\quad \times (298.15 \text{ K})(\ln K) \\ K &= 285.64 \end{aligned}$$

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023 4.0 points

A mixture of $PCl_5(g)$ and $Cl_2(g)$ is placed into a closed container. At equilibrium it is found that $[PCl_5] = 0.72$ M, $[Cl_2] = 0.45$ M and $[PCl_3] = 0.12$ M.

$$PCl_5 \rightleftharpoons PCl_3 + Cl_2$$

What is the value of K_c for the reaction?

1. 181

2. 0.075 **correct**

3. 0.0375

4. 0.225

5. 0.15

Explanation:
$[PCl_5] = 0.72$ M $\quad [Cl_2] = 0.45$ M
$[PCl_3] = 0.12$ M

$$\begin{aligned} K_c &= \frac{[Cl_2]\,[PCl_3]}{[PCl_5]} = \frac{(0.45 \text{ M})(0.12 \text{ M})}{0.72 \text{ M}} \\ &= 0.075 \text{ M} \end{aligned}$$

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024 4.0 points

$K_c = 2.6 \times 10^8$ at 825 K for the reaction

$$2\,H_2(g) + S_2(g) \rightleftharpoons 2\,H_2S(g)$$

The equilibrium concentration of H_2 is 0.0020 M and that of S_2 is 0.0010 M. What is the equilibrium concentration of H_2S?

1. 0.10 M

2. 1.02 M **correct**

3. 0.0010 M

4. 10 M

Explanation:
$K_c = 2.6 \times 10^8 \quad [H_2]_{eq} = 0.0020$ M
$[S_2]_{eq} = 0.0010$ M

$$2\,H_2(g) + S_2 \rightleftharpoons 2\,H_2S$$

$$K_c = \frac{[H_2S]^2}{[H_2]^2\,[S_2]}$$

$$\begin{aligned} [H_2S] &= \sqrt{K_c\,[H_2]^2\,[S_2]} \\ &= \sqrt{(2.6 \times 10^8)\,(0.0020 \text{ M})^2\,(0.0010 \text{ M})} \\ &= 1.0 \text{ M} \end{aligned}$$

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025 4.0 points

A 2.000 liter vessel is filled with 4.000 moles of SO_3 and 6.000 moles of O_2. When the reaction

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

comes to equilibrium a measurement shows that only 1.000 mole of SO_3 remains. How many moles of O_2 are in the vessel at equilibrium?

1. None of these is correct.

2. 7.000 mol

3. 12.000 mol

4. 3.750 mol

5. 7.500 mol **correct**

Explanation:

Initially,

$$[SO_3] = \frac{4\text{ mol}}{2\text{ L}} = 2\text{ M} \qquad [O_2] = \frac{6\text{ mol}}{2\text{ L}} = 3\text{ M}$$

	$2\,SO_3$ (g) $\rightleftharpoons$	$2\,SO_2$ (g) +	O_2 (g)
ini, M	2	0	3
Δ, M	$-2x$	$2x$	x
eq, M	$2-2x$	$2x$	$3+x$

At equilibrium,

$$[SO_3]_{eq} = \frac{1\text{ mol}}{2\text{ L}} = 0.5\text{ M, so}$$

$$2-2x = 0.5$$
$$-2x = -1.5$$
$$x = 0.75$$

Thus

$$[O_2] = 3 + x = 3.75\text{ M}$$

$$\text{mol } O_2 = (3.75\text{ mol/L})\,(2\text{ L}) = 7.5\text{ mol}\,.$$

026 4.0 points

At $T = 500°\text{C}$, $K_c = 36$ for the gas-phase reaction

$$A + B \rightleftharpoons C + D$$

Starting with 2.04 moles each of A and B in a 5.00 liter container, what will be the equilibrium concentration of C at this temperature?

1. 0.349714
2. 0.6516
3. 0.4896
4. 0.675692
5. 0.315333
6. 0.256
7. 0.600889
8. 0.607286
9. 0.452
10. 0.4788

Correct answer: 0.349714 M.

Explanation:

$$[A] = \frac{2.04\text{ mol}}{5\text{ L}} = 0.408\text{ M} \qquad T = 500°\text{C}$$

$$[B] = \frac{2.04\text{ mol}}{5\text{ L}} = 0.408\text{ M} \qquad K_c = 36$$

	A +	B $\rightleftharpoons$	C +	D
ini, M	0.408	0.408	0	0
Δ, M	$-x$	$-x$	x	x
eq, M	$0.408-x$	$0.408-x$	x	x

$$\frac{[C][D]}{[A][B]} = 36$$

$$\frac{x^2}{(0.408-x)^2} = 36$$

$$\frac{x}{0.408-x} = 6$$

$$x = 2.448 - 6\,x$$

$$x = [C] = 0.349714\text{ M}$$

027 4.0 points

Suppose the reaction

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

has an equilibrium constant $K_c = 49$ and the initial concentration of H_2 and I_2 is 0.5 M and HI is 0.0 M. Which of the following is the correct value for the final concentration of HI(g)?

1. 0.219 M

2. 0.778 M **correct**

3. 0.389 M

4. 0.599 M

5. 0.250 M

Explanation:

$K_c = 49$ $[H_2]_{ini} = 0.5$ M

$[I_2]_{ini} = 0.5$ M $[HI]_{ini} = 0$ M

	$H_2(g)$	+	$I_2(g)$	$\rightleftharpoons 2\,HI(g)$
Ini, M	0.5		0.5	0
Δ, M	$-x$		$-x$	$+2\,x$
Equil, M	$0.5 - x$		$0.5 - x$	$2\,x$

$$\begin{aligned}
K_c &= \frac{[HI]^2}{[H_2]\,[I_2]} \\
49 &= \frac{(2x)^2}{(0.5 - x)^2} \\
7 &= \frac{2x}{0.5 - x} \\
7(0.5 - x) &= 2x \\
3.5 - 7x &= 2x \\
3.5 &= 9x \\
x &= \frac{3.5}{9} = 0.389 \text{ M}
\end{aligned}$$

Looking back at our equilibrium values, we see that the final concentration of HI is equal to $2x$, so $2(0.389) = 0.778$ M.