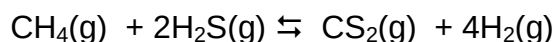


Equilibrium Worksheet Key

1. $2\text{NH}_3(\text{g}) \rightleftharpoons \text{N}_2(\text{g}) + 3\text{H}_2(\text{g})$ At 500 K, the following concentrations were measured: $[\text{N}_2] = 3.0 \times 10^{-2} \text{ M}$, $[\text{H}_2] = 3.7 \times 10^{-2} \text{ M}$, $[\text{NH}_3] = 1.6 \times 10^{-2} \text{ M}$. What is K_c ?

$$K_c = \frac{[\text{N}_2][\text{H}_2]^3}{[\text{NH}_3]^2} = \frac{(3.0 \times 10^{-2})(3.7 \times 10^{-2})^3}{(1.6 \times 10^{-2})^2} = 5.9 \times 10^{-3}$$

2. At 1000 K, the equilibrium partial pressures for the reaction below are: $\text{CH}_4 = 0.20 \text{ atm}$, $\text{H}_2\text{S} = 0.25 \text{ atm}$, $\text{CS}_2 = 0.52 \text{ atm}$, and $\text{H}_2 = 0.10 \text{ atm}$. What is K_p ?



$$K_p = \frac{(P_{\text{CS}_2}) \cdot (P_{\text{H}_2})^4}{(P_{\text{CH}_4}) \cdot (P_{\text{H}_2\text{S}})^2} = \frac{(0.52)(0.10)^4}{(0.20)(0.25)^2} = 4.2 \times 10^{-3}$$

3. $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ At 375 °C, $K_c = 2.79 \times 10^{-5}$.

a) What is K_p ? $K_p = K_c(RT)^{\Delta n}$ where $R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}$, $T = 375 + 273 = 648 \text{ K}$

$$\Delta n = \# \text{ product gas moles} - \# \text{ reactant gas moles} = 2 - 4 = -2$$

$$K_p = 2.79 \times 10^{-5} (0.08206 \times 648)^{-2} = \frac{2.79 \times 10^{-5}}{(53.175)^2} = 9.87 \times 10^{-9}$$

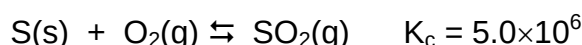
b) What is P_{NH_3} if $P_{\text{H}_2} = 1.24 \text{ atm}$ and $P_{\text{N}_2} = 2.17 \text{ atm}$ at equilibrium?

$$K_p = \frac{(P_{\text{NH}_3})^2}{(P_{\text{N}_2}) \cdot (P_{\text{H}_2})^3} \Rightarrow 9.87 \times 10^{-9} = \frac{(P_{\text{NH}_3})^2}{(2.17)(1.24)^3}$$

$$(P_{\text{NH}_3})^2 = 9.87 \times 10^{-9} (2.17)(1.24)^3 = 4.0836 \times 10^{-8}$$

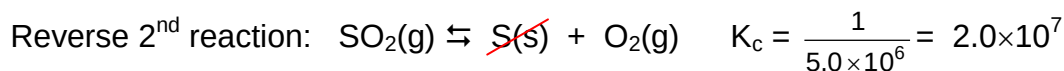
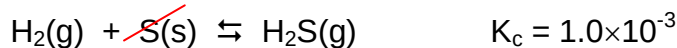
$$P_{\text{NH}_3} = \sqrt{4.0836 \times 10^{-8}} = 2.02 \times 10^{-4} \text{ atm}$$

4. Given the equations: $\text{H}_2(\text{g}) + \text{S}(\text{s}) \rightleftharpoons \text{H}_2\text{S}(\text{g})$ $K_c = 1.0 \times 10^{-3}$



Calculate the value of K_c for $\text{H}_2(\text{g}) + \text{SO}_2(\text{g}) \rightleftharpoons \text{H}_2\text{S}(\text{g}) + \text{O}_2(\text{g})$

- How can we combine the reactions to match the last reaction?
- Need to reverse the second reaction, then add the two reactions together:

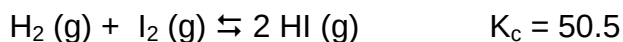


5. For the reaction, $\text{B} \rightleftharpoons 2\text{A}$, $K_c = 2$. Suppose 3.0 moles of A and 3.0 moles of B are introduced into a 2.00 L flask.

$$[\text{A}] = \frac{3.0 \text{ mol}}{2.00 \text{ L}} = 1.5 \text{ M} \quad [\text{B}] = \frac{3.0 \text{ mol}}{2.00 \text{ L}} = 1.5 \text{ M} \quad Q = \frac{[\text{A}]^2}{[\text{B}]} = \frac{(1.5)^2}{(1.5)} = 1.5$$

- a) Is this system at equilibrium? **No**, **Q is less than K so not at eq**
- b) In which direction will the reaction proceed to reach equilibrium? **Q < K so there are not enough products** $\Rightarrow$ **Reaction shifts right (forwards towards products)**
- c) As the system moves towards equilibrium, what happens to the concentration of B? A?
B ↓ and A ↑ (Reactants ↓ and products ↑ as rxn shifts $\rightarrow$)

6. a) Calculate the equilibrium concentrations of all species for the following reaction if the initial concentrations of H_2 and I_2 are both 1.00 M.



- **Steps: Set up ICE table and K_c , plug Eq terms into K_c & solve for x**

	$H_2(g)$	+	$I_2(g)$	$\rightleftharpoons$	$2HI(g)$
I, M	1.00		1.00		0
C, M	-x		-x		+2x
E, M	1.00 - x		1.00 - x		2x

$$K_c = \frac{[HI]^2}{[H_2][I_2]} \quad 50.5 = \frac{(2x)^2}{(1.00-x)(1.00-x)} = \frac{(2x)^2}{(1.00-x)^2} \quad \leftarrow \text{Perfect square}$$

$$\text{Take square root both sides: } 7.106 = \frac{2x}{1.00-x}$$

$$7.106(1.00 - x) = 2x \quad \Rightarrow \quad 7.106 - 7.106x = 2x \quad \Rightarrow \quad 7.106 = 9.106x$$

$$x = \frac{7.106}{9.106} = 0.780 \quad \text{thus } x = \mathbf{0.0780 \text{ M}}$$

$$[HI] = 2x = 2(0.780 \text{ M}) = 1.56 \text{ M}; [H_2] = [I_2] = 1.00 - x = 1.00 - 0.780 = 0.22 \text{ M}$$

Eq concs: **[HI] = 1.56 M, [H₂] = [I₂] = 0.22 M**

- b) For the reaction above, if $K_p = 50.5$ and the initial pressures are $HI = 0.975 \text{ atm}$, $H_2 = 0.105 \text{ atm}$ and $I_2 = 0.105 \text{ atm}$, what are the equilibrium pressures for all the substances?

- **Steps: Given initial pressures for all substances so calculate Q to see if at equilibrium. If Q <, rxn shifts $\rightarrow$, If Q >, rxn shifts $\leftarrow$**
- **Set up ICE table (+ signs on side where substances $\uparrow$, - signs on opposite side)**
- **Plug Eq terms into K_p & solve for x**

$$Q = \frac{(P_{HI})^2}{(P_{H_2})(P_{I_2})} \quad Q = \frac{(0.975)^2}{(0.105)(0.105)} = 86.2$$

$Q > K_p$ so too many products are present & reaction shifts left to attain eq
 Rxn shifts $\leftarrow$, thus products decrease (- signs) and reactants increase (+ signs)

	$H_2(g)$	+	$I_2(g)$	$\rightleftharpoons$	$2HI(g)$
I, atm	0.105		0.105		0.975
C, atm	+x		+x		-2x
E, atm	0.105 + x		0.105 + x		0.975 - 2x

$$Q = \frac{(P_{HI})^2}{(P_{H_2}) \cdot (P_{I_2})} \quad 50.5 = \frac{(0.975-2x)^2}{(0.105+x)(0.105+x)} = \frac{(0.975-2x)^2}{(0.105+x)^2} \quad \Leftarrow \text{Perfect square}$$

Take square root both sides: $7.106 = \frac{0.975-2x}{0.105+x}$

$$7.106 (0.105 + x) = 0.975 - 2x$$

$$0.74613 + 7.106x = 0.975 - 2x$$

$$9.106x = 0.22887$$

$$x = \frac{0.22887}{9.106} = 0.0251 \quad \text{thus } x = 0.0251 \text{ atm}$$

$$P_{HI} = 0.975 - 2x = 0.975 - 2(0.0251) = 0.925 \text{ atm};$$

$$P_{H_2} = P_{I_2} = 0.105 + x = 0.105 + 0.0251 = 0.130 \text{ atm}$$

Eq pressures: $P_{HI} = 0.925 \text{ atm}, P_{H_2} = P_{I_2} = 0.130 \text{ atm}$

6. Calculate the equilibrium concentrations for the reaction below if the initial $[N_2] = 0.80 \text{ M}$ and the initial $[O_2] = .20 \text{ M}$



Set up ICE and plug eq concs into K_c !

	$N_2(g)$	+	$O_2(g)$	$\rightleftharpoons$	$2 NO(g)$
I, M	0.80		0.20		0
C, M	-x		-x		+2x
E, M	$0.80 - x$		$0.20 - x$		+2x

$$K_c = \frac{[NO]^2}{[N_2][O_2]} \quad 1.0 \times 10^{-5} = \frac{(2x)^2}{(0.80-x)(0.20-x)}$$

- Assume x is small (its much smaller than 0.80 or 0.20) because K_c is small (less than 10^{-3})

$$1.0 \times 10^{-5} = \frac{(2x)^2}{(0.80)(0.20)}$$

$$(2x)^2 = 1.0 \times 10^{-5} (0.80)(0.20)$$

$$4x^2 = 1.6 \times 10^{-6}$$

$$x = \sqrt{\frac{1.6 \times 10^{-6}}{4}} = 6.3 \times 10^{-4} \text{ M}$$

$$[N_2] = 0.80 - 6.3 \times 10^{-4} = 0.80 \text{ M}, [O_2] = 0.20 \text{ M} - 6.3 \times 10^{-4} = 0.20 \text{ M},$$

$$[NO] = 2(6.3 \times 10^{-4} \text{ M}) = 1.3 \times 10^{-3} \text{ M}$$

- Check assumption: $(x / \text{smallest initial concentration}) * 100\% < 5\%$
- Check answer by plugging concentrations into K_c

7. Calculate the equilibrium concentrations of all species if 3.000 moles of H_2 and 6.000 moles of F_2 are placed in a 3.000 L container.



- Note: Different initial amounts and large K so short cuts don't apply!**

$$[\text{H}_2] = \frac{3.000 \text{ moles}}{3.000 \text{ L}} = 1.000 \text{ M}$$

$$[\text{F}_2] = \frac{6.000 \text{ moles}}{3.000 \text{ L}} = 2.000 \text{ M}$$

	$\text{H}_2(\text{g})$	+	$\text{F}_2(\text{g})$	$\rightleftharpoons$	$2\text{HF}(\text{g})$
I, M	1.000		2.000		0
C, M	-x		-x		+2x
E, M	1.000 - x		2.000 - x		+2x

$$K_c = \frac{[\text{HF}]^2}{[\text{H}_2][\text{F}_2]} \quad 1.15 \times 10^2 = \frac{(2x)^2}{(1.000 - x)(2.000 - x)}$$

$$(1.15 \times 10^2)(1.000 - x)(2.000 - x) = 4x^2$$

$$(1.15 \times 10^2)(2.000 - 1.000x - 2.000x + x^2) = 4x^2$$

$$(1.15 \times 10^2)(2.000 - 3.000x + x^2) = 4x^2$$

$$2.30 \times 10^2 - 3.45 \times 10^2 x + 1.15 \times 10^2 x^2 = 4x^2$$

$$(1.11 \times 10^2)x^2 - (3.45 \times 10^2)x + 2.30 \times 10^2 = 0$$

$$ax^2 + bx + c = 0 \quad x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = \frac{-(-3.45 \times 10^2) \pm \sqrt{(-3.45 \times 10^2)^2 - 4(1.11 \times 10^2)(2.30 \times 10^2)}}{2 \times 1.11 \times 10^2}$$

$$x = 2.14 \text{ or } 0.968$$

- Note: x can't be 2.14 because this would give us negative equilibrium concentrations which are physically meaningless, so x = 0.968.**

$$[\text{H}_2] = 1.000 - 0.968 = 3.2 \times 10^{-2} \text{ M}; [\text{F}_2] = 2.000 - 0.968 = 1.032 \text{ M}; [\text{HF}] = 2(0.968) = 1.936 \text{ M}$$