

Name  
Kc/Kp Multiple choice

d 1. (ebbing14.3)  
The equilibrium expression for  $K_c$  for the system  
 $\text{CO}_{2(g)} + \text{CaO}_{(s)} \rightarrow \text{CaCO}_{3(s)}$  is

- a.  $[\text{CaCO}_3]/[\text{CO}_2][\text{CaO}]$   
 b.  $[\text{CaCO}_3]/[\text{CO}_2]$   
 c.  $[\text{CO}_2]$   
 d.  $1/[\text{CO}_2]$   
 e.  $[\text{CO}_2][\text{CaO}]$

e 2. (Ebbing 14.4)  
In which of the following does the reaction go the least to completion (see the following K values)

a.  $10E5$   
 b.  $10E3$   
 c.  $10E0$   
 d.  $10E-3$   
 e.  $10E-5$  *↳ smallest K*

a 3. (Ebbing 14.7)  
Carbon disulfide and chlorine react according to the following equation:  
 $\text{CS}_{2(g)} + 3\text{Cl}_{2(g)} \leftrightarrow \text{S}_2\text{Cl}_{2(g)} + \text{CCl}_4(g)$   
When 1.00 mol of  $\text{CS}_2$  and 3.00 mol of  $\text{Cl}_2$  are placed in a 2.00L container and allowed to come to equilibrium, the mixture is found to contain 0.250 mol of  $\text{CCl}_4$ . What is the amount of  $\text{Cl}_2$  at equilibrium?

- a. 2.25 mol  
 b. 2.75 mol  
 c. 0.75 mol  
 d. .25 mol  
 e. .50 mol
- ICE table:*  
 $\text{CS}_{2(g)} + 3\text{Cl}_{2(g)} \rightleftharpoons \text{S}_2\text{Cl}_{2(g)} + \text{CCl}_4(g)$   
 I 1.00 3.00 0 0  
 C -0.25 -0.750 +0.250 +0.250  
 E 0.75 2.250 0.250 0.250  
 mol

d 5. (Ebbing14.11)  
Which expression correctly describes the equilibrium constant for the following reaction?  
 $4\text{NH}_{3(g)} + 5\text{O}_2 \leftrightarrow 4\text{NO}_{(g)} + 6\text{H}_2\text{O}(g)$

a.  $K_c = 4[\text{NH}_3] + 5[\text{O}_2]/6[\text{H}_2\text{O}] + 4[\text{NO}]$   
 b.  $K_c = 6[\text{H}_2\text{O}] + 4[\text{NO}]/4[\text{NH}_3] + 5[\text{O}_2]$   
 c.  $K_c = [\text{H}_2\text{O}][\text{NO}]/[\text{NH}_3][\text{O}_2]$   
 d.  $K_c = [\text{H}_2\text{O}]^6[\text{NO}]^4/[\text{NH}_3]^4[\text{O}_2]^5$   
 e.  $K_c = [\text{NH}_3]^4[\text{O}_2]^5/[\text{H}_2\text{O}]^6[\text{NO}]^4$

$$K_c = \frac{[\text{NO}]^4 [\text{H}_2\text{O}]^6}{[\text{NH}_3]^4 [\text{O}_2]^5}$$

b 7. (ebbing14.19)  
Consider the reaction system  
 $\text{Br}_2(g) + \text{Cl}_2(g) \leftrightarrow 2\text{BrCl}(g)$

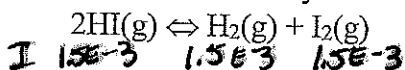
$$K_c = \frac{[\text{BrCl}]^2}{[\text{Br}_2][\text{Cl}_2]} = \frac{(0.015)^2}{(0.006)(0.0095)} = 3.947$$

At a given temperature. When the system is at equilibrium, the molar concentrations of  $\text{Br}_2$ ,  $\text{Cl}_2$  and  $\text{BrCl}$  are 0.0060M, 0.0095M, and 0.015M, respectively. The value of  $K_c$  for this system is

- a. .025  
 b. 3.9  
 c. 27  
 d. 53  
 e. 260

d 8. (ebbing14.21)

For the reaction system



$K_c = 0.020$  at 720K. If the initial concentrations of HI,  $\text{H}_2$ , and  $\text{I}_2$  are all  $1.50\text{E-}3\text{M}$  at 720K, which one of the following statements is correct?

$K = 0.02$   $< Q = 1$   
 reactant favored  
 more reactants

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = 0.020 \quad Q = \frac{(1.5\text{E-}3)(1.5\text{E-}3)}{(1.5\text{E-}3)^2} = 1$$

a. The system is at equilibrium.

no  $K \neq Q$

d. The concentration of HI will increase as the system is approaching equilibrium.

more reactants

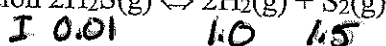
b. The concentrations of HI and  $\text{I}_2$  will increase as the system is approaching equilibrium.

decrease

e. The concentrations of  $\text{H}_2$  and  $\text{I}_2$  will increase as the system is approaching equilibrium. (decrease)

c. The concentrations of  $\text{H}_2$  and HI will decrease as the system approaches equilibrium.

b 9. (ebbing14.22) For the reaction  $2\text{H}_2\text{S}(\text{g}) \rightleftharpoons 2\text{H}_2(\text{g}) + \text{S}_2(\text{g})$



$$Q = \frac{[\text{H}_2]^2[\text{S}_2]}{[\text{H}_2\text{S}]^2} = \frac{(1.5)(1)^2}{(0.01)^2} = 15000$$

at a certain temperature  $K_c$  equals 4500. What will happen when 0.010 mol of  $\text{H}_2\text{S}(\text{g})$ , 1.0 mol of  $\text{H}_2$  and 1.5 mol of  $\text{S}_2$  are added to a 2.0L container and the system is brought to the temperature at which  $K_c = 4500$ .

more product

$Q > K_c$   
15,000 4500

less products + more reactant  $\text{H}_2\text{S}$

Shift ←

a. Nothing, the system is at equilibrium.

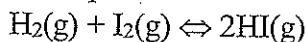
d. More  $\text{S}_2$  will be formed than  $\text{H}_2$ .

b. More  $\text{H}_2\text{S}$  will be formed.

e. The amount of  $\text{H}_2$  formed will be half the amount of  $\text{S}_2$ .

c. More  $\text{H}_2$  will be formed than  $\text{S}_2$

d 10. (ebbing14.23) A 1.00 mol sample of HI is placed in a 1-L vessel at 460C, and the reaction system is allowed to come to equilibrium. The HI partially decomposes, forming 0.11 mol  $\text{H}_2$  and 0.11 mol  $\text{I}_2$ . What is the equilibrium constant for the reaction?



a. 0.020

b. 7.1

c. 8.1

d. 50

e. 65

$$\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$$

I	0.11	0.11	1.00
S	+0.11	+0.11	-0.22
E	0.11	0.11	0.78

$$K = \frac{(0.78)^2}{(0.11)(0.11)} = 50$$

11. (ebbing14.25)

Consider the equilibrium  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$  at a certain temperature. An equilibrium mixture in an 8.00L vessel contains .800 mol  $\text{N}_2$ , and 1.20 mol  $\text{H}_2$ . What is the value of  $K_c$ ?

a. 1.85

b. 29.6

c. 37.4

d. 75.8

e. 119

need more information

$$\frac{0.8}{8} = 0.1\text{M}$$

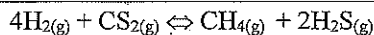
$$\frac{1.2}{8} = 0.15$$

mols in 1L

$$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$$

I			
S			
E	0.1	0.15	

C 12.



The system reaches equilibrium according to the equation above. A mixture of 2.5 mol H<sub>2</sub>, 1.50 mol CS<sub>2</sub>, 1.50 mol CH<sub>4</sub> and 2.00 mol H<sub>2</sub>S is placed in a 5L reaction vessel. When equilibrium is achieved, the concentration of CH<sub>4</sub> has become 0.25M.

Changes in concentration occur as this system approaches equilibrium. Which expression gives the best comparison of the changes in those concentration shown in the ratio below?

$$\frac{\Delta[\text{H}_2\text{S}]}{\Delta[\text{CS}_2]}$$

- a. +2/+1
- b. +2/-1
- c. -2/+1
- d. -1/+1
- e. -1/+2

Handwritten ICE table for the reaction:

	$4\text{H}_2(\text{g})$	$+\text{CS}_2(\text{g})$	$\rightleftharpoons$	$\text{CH}_4(\text{g})$	$+2\text{H}_2\text{S}(\text{g})$
I	2.5	1.5		1.5	2.0
S	$-4x$	$-x$		$+x$	$+2x$
E	0.7	0.35		0.25	0.4

Additional handwritten notes:

- $\Delta[\text{H}_2\text{S}] = -2x$
- $\Delta[\text{CS}_2] = +x$
- 2.5 mol / 5L = 0.5 M
- X = 0.05 mol in 1L
- Change in M is  $-2x$
- $2(0.05) = 0.1\text{M}$
- $0.1\text{M} = \frac{x \text{ mol}}{5\text{L}}$
- $x = 0.5 \text{ mol}$

b 13.

What is the change in the number of moles of H<sub>2</sub>S(g) present as the system moves from its original state to the equilibrium described?

- a. -1.25
- b. -0.50
- c. -0.25
- d. +0.25
- e. +0.50

Handwritten calculations for question 13:

- $2(0.05) = 0.1\text{M}$
- $0.1\text{M} = \frac{x \text{ mol}}{5\text{L}}$
- $x = 0.5 \text{ mol}$

e 14.

When equilibrium is achieved, the concentration of CH<sub>4</sub> has become 0.25M.

What is the number of moles of CS<sub>2</sub>(g) at equilibrium?

- a. .25
- b. .35
- c. .75
- d. 1.25
- e. 1.75

Handwritten calculation for question 14:

- $0.35\text{M} = \frac{x}{5\text{L}}$
- $x = 1.75 \text{ mol}$

b 15.

When equilibrium is achieved, the concentration of CH<sub>4</sub> has become 0.25M.

What is the concentration in moles per liter of H<sub>2</sub> at equilibrium?

- a. 5
- b. .70
- c. 1.00
- d. 2.5
- e. 3.5

M

Handwritten ICE table for the reaction:

	$4\text{H}_2(\text{g})$	$+\text{CS}_2(\text{g})$	$\rightleftharpoons$	$\text{CH}_4(\text{g})$	$+2\text{H}_2\text{S}(\text{g})$
I	2.5	1.50		1.50	2.00
S	$-4.0$	$-0.25$		$+0.25$	$+0.50$
E	3.5	1.25		1.25	1.50

Additional handwritten notes:

- total moles
- $\frac{3.5 \text{ mol}}{5\text{L}} = 0.7\text{M}$
- $0.25\text{M} = \frac{x}{5\text{L}}$
- $x = 1.25 \text{ mols}$