

Quantitative Equilibrium Problems

Note ①, ②, ③ etc shows the order of figuring

1. a)

	C	D	⇌	C	+	D
Initial []	0.200 mol/L					
Initial Amount	0.200 mol					
Final Amount	0.200 - 0.030 ^① = 0.170 ^② mol			0.030 ^③ mol		0.030 ^③ mol
Final []	0.170 ^④ mol/L			0.030 ^⑤ mol/L		0.030 ^⑥ mol/L

assume 1L volume

b) $K_{eq} = \frac{[C][D]}{[CD]}$

$K_{eq} = \frac{(0.030)(0.030)}{(0.170)}$

$K_{eq} = 0.00529$

* $0.200 \times \frac{15\%}{100\%} = 0.030$

2. a)

	A	+	B	⇌	AB
Initial []	/		/		/
Initial Amount	1.00 mol		1.00 mol		∅
Final Amount	1.00 - 0.40 ^① = 0.60 ^② mol		1.00 - 0.40 ^③ = 0.60 ^④ mol		0.40 ^⑤ mol
Final []	0.60 ^⑥ mol/L		0.60 ^⑦ mol/L		0.40 ^⑧ mol/L*

1L

* $\frac{\text{kmol}}{\text{m}^3} = \frac{\text{mol}}{\text{L}}$

$$\begin{aligned} \% \text{ A convert} &= \frac{\text{converted A}}{\text{total A}} \times 100\% \\ &= \frac{0.40}{1.00} \times 100\% \\ &= 40\% \end{aligned}$$



Initial []	/	/	/	1.0 L
Initial amount	0.50 mol	0.50 mol	∅	
Final amount	0.50 - 0.0125 ^② = 0.4875 mol ^③	0.50 - 0.0125 ^② = 0.4875 mol ^③	0.025 mol ^①	
Final []	0.4875 mol/L ^④	0.4875 mol/L ^④	0.025 mol/L	

$$K_{eq} = \frac{[XY]^2}{[X_2][Y_2]}$$

$$= \frac{(0.025)^2}{(0.4875)(0.4875)}$$

$$= 0.00263$$



Initial []	/	/	/	4.0 L
Initial amount	1.00 mol	∅	∅	
Final amount	1.00 - 0.200 ^② 0.800 mol ^③	0.400 mol ^②	0.200 mol ^①	
Final []	0.200 mol/L ^④	0.100 mol/L ^④	0.050 mol/L	

$$K_{eq} = \frac{[B]^2 [C]}{[A]}$$

$$K_{eq} = \frac{(0.100)^2 (0.050)}{(0.200)}$$

$$K_{eq} = 0.0025$$

5.



5.00L

 $K_{eq} = 3.0$

@ equilibrium

Initial []	/	/	/
Initial Amount	∅	1.00 mol	1.00 mol
Final Amount	2x	1-x	1-x
Final []	$\frac{2x}{5}$	$\frac{1-x}{5}$	$\frac{1-x}{5}$

Let x represent the amount of Y or Z that reacts

$$K_{eq} = \frac{[Y][Z]}{[X]^2}$$

$$3.0 = \frac{\left(\frac{1-x}{5}\right)\left(\frac{1-x}{5}\right)}{\left(\frac{2x}{5}\right)^2}$$

$$3.0 = \left(\frac{1-x}{5}\right)\left(\frac{1-x}{5}\right)\left(\frac{5}{2x}\right)\left(\frac{5}{2x}\right)$$

$$12x^2 = x^2 - 2x + 1$$

$$0 = 11x^2 + 2x - 1$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = \frac{-2 \pm \sqrt{2^2 - 4(11)(-1)}}{2(11)}$$

$$x = \frac{-2 \pm \sqrt{48}}{22}$$

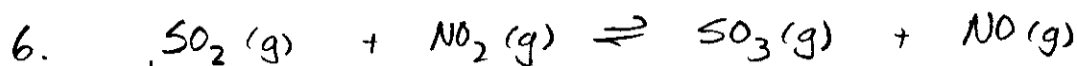
$\therefore x = -0.406$ or $x = 0.224$
extraneous ($x \neq -$)

← amount of X

$$n_X = 2x$$

$$= 2(0.224)$$

$$= 0.448 \text{ mol}$$

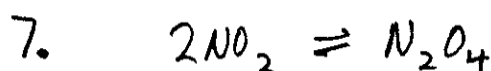


amounts	0.400 mol	0.0500 mol	0.300 mol	0.200 mol	1.00 L
[]	0.400 mol/L	0.0500 mol/L	0.300 mol/L	0.200 mol/L	

$$K_{\text{eq}} = \frac{[\text{SO}_3][\text{NO}]}{[\text{SO}_2][\text{NO}_2]}$$

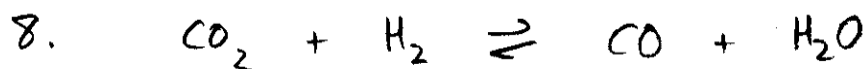
$$K_{\text{eq}} = \frac{(0.300)(0.200)}{(0.400)(0.0500)}$$

$$K_{\text{eq}} = 3.00$$



$$K_{\text{eq}} = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2}$$

$$\begin{aligned} [\text{N}_2\text{O}_4] &= K_{\text{eq}} [\text{NO}_2]^2 \quad \text{kmol/m}^3 = \text{mol/L} \\ &= 1.15 (0.50)^2 \\ &= 0.288 \end{aligned}$$



$$K_{\text{eq}} = \frac{[\text{CO}][\text{H}_2\text{O}]}{[\text{CO}_2][\text{H}_2]}$$

$$K_{\text{eq}} = \frac{(0.00133)(0.00133)}{(0.00117)(0.00117)}$$

$$K_{\text{eq}} = 1.29$$

9.

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

a)+b) Initial []	/	/	/
Initial Amount	1 mol	∅	∅
Final amount	1 - 0.200 = 0.800 mol	0.100 mol	0.300 mol
Final []	0.800 mol/L	0.100 mol/L	0.300 mol/L

c) 1L $K_{eq} = \frac{[\text{N}_2][\text{H}_2]^3}{[\text{NH}_3]^2}$

$$K_{eq} = \frac{(0.100)(0.300)^3}{(0.800)^2}$$

$$K_{eq} = 0.00422$$

d) S: ↑ [H₂] H: use H₂
 R: ↓ [H₂] O: shift left

e)+f) only a change in temp will affect the value of K_{eq}

10.

$$\text{CO}_2 + \text{H}_2 \rightleftharpoons \text{H}_2\text{O} + \text{CO}$$

a)	/	/	/	/
	0.5 mol	0.5 mol	∅	∅
	0.5 - x	0.5 - x	x	x
	0.5 - x	0.5 - x	x	x

1L

$$K_{eq} = 2.00$$

$$\rightarrow x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = \frac{2 \pm \sqrt{(-2)^2 - 4(1)(0.50)}}{2(1)}$$

$$x = \frac{2 \pm 1.414}{2}$$

$$x = \cancel{1.707} \text{ or } x = 0.293$$

extraneous

$$0.5 - x \neq 0$$

$$\therefore x \leq 0.5$$

$$\therefore [\text{CO}_2] = 0.5 - x$$

$$= 0.5 - 0.293$$

$$= 0.207 \text{ mol/L}$$

$$\therefore [\text{H}_2] = 0.207 \text{ mol/L}$$

$$[\text{H}_2\text{O}] = 0.293 \text{ mol/L}$$

$$[\text{CO}] = 0.293 \text{ mol/L}$$

Let x represent the amount of H₂O that forms

$$K_{eq} = \frac{[\text{H}_2\text{O}][\text{CO}]}{[\text{CO}_2][\text{H}_2]}$$

$$2.00 = \frac{(x)(x)}{(0.5-x)(0.5-x)}$$

$$2(x^2 - x + 0.25) = x^2$$

$$2x^2 - 2x + 0.50 = x^2$$

$$0 = x^2 - 2x + 0.50$$

b) given the symmetrical aspects of this equilibrium, same concentrations would be achieved

11. a)

	H_2	$+ CO_2$	\rightleftharpoons	H_2O	$+ CO$	
Amounts	1.17 mol	1.17 mol		1.33 mol	1.33 mol	10.0L
[]s	0.117 mol	0.117 mol		0.133 mol	0.133 mol	

$$K_{eq} = \frac{[H_2O][CO]}{[H_2][CO_2]}$$

$$K_{eq} = \frac{(0.133)(0.133)}{(0.117)(0.117)}$$

$$K_{eq} = 1.29$$

b) S: $\uparrow V \Rightarrow \downarrow P$

R: $\uparrow P$

H: make more moles of gas

D: no shift (equi-molar)

E: no change in n_{H_2O}

c) $\downarrow [H_2O]$ $\downarrow C = \frac{n}{V}$ if V increases, C decreases

d) no effect, only temperature affects K_{eq}

e) $H_2 + CO_2 \rightleftharpoons H_2O + CO$

Initial []	/	/	/	/
Initial Amount	1.17 mol	1.17 mol	$(1.33 + x)$ mol	1.33 mol
Final Amount	$1.17 + 0.33^*$ $= 1.50$ mol	$1.17 + 0.33$ $= 1.50$ mol	$1.33 + x - 0.33$ $1.00 + x$ mol	$1.33 - 0.33$ 1.00 mol
Final []	0.150 mol/L	0.150 mol/L	$\frac{1.00 + x}{10}$ mol/L	0.100 mol/L

Let x represent the amount of H_2O injected

* note that the shift must 0.33

$$K_{eq} = \frac{[H_2O][CO]}{[H_2][CO_2]}$$

$$1.29 = \frac{\left(\frac{1.00 + x}{10}\right)(0.100)}{(0.150)(0.150)}$$

$$0.0290 = \left(\frac{1.00 + x}{10}\right)(0.100)$$

$$2.90 = 1.00 + x$$

$$x = 1.9025 \text{ mol } H_2O \text{ injected}$$