Quantitative Equilibrium Problems

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

1. a) $CD \rightleftharpoons C + D$ assume 1L volume

	CD	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 ④ mol/L	0.030 ④ mol/L

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

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

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

$K_{eq} = 0.00529$

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2. a) $A + B \rightleftharpoons AB$ 1L

	A	B	AB
Initial [ ]	/	/	/
Initial amount	1.00 mol	1.00 mol	∅
Final amount	1.00 - 0.40 ③ = 0.60 ④ mol	1.00 - 0.40 ③ = 0.60 ④	0.40 ② mol
Final [ ]	0.60 ⑤ mol/L	0.60 ⑤ mol/L	0.40 ① mol/L *

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

$\%\ A\ convert = \frac{converted\ A}{total\ A} \times 100\%$

$= \frac{0.40}{1.00} \times 100\%$

$= 40\%$

3. $X_2 + Y_2 \rightleftharpoons 2XY$

	X_2	Y_2	$2XY$
Initial [ ]	/	/	/
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

1.0 L

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

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

$$= 0.00263$$

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4. $A \rightleftharpoons 2B + C$

	A	2B	C
Initial [ ]	/	/	/
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

4.0 L

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

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

$$K_{eq} = 0.0025$$

5. $2X \rightleftharpoons Y + Z$

5.00 L

$K_{eq} = 3.0$

	$2X$	Y	Z
Initial [ ]	/	/	/
Initial amount	∅	1.00 mol	1.00 mol
Final amount	$2x$	$1-x$	$1-x$
Final [ ]	$\frac{2x}{5}$	$\frac{1-x}{5}$	$\frac{1-x}{5}$

@ equilibrium

Let x represent the amount of Y or Z that reacts

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

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

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

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

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

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

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

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

$\therefore$ $\cancel{x = -0.406}$ or $x = 0.224$

extraneous

$(x \neq -)$

amount of X

$$n_X = 2x$$
$$= 2(0.224)$$
$$= 0.448 \text{ mol}$$

6. $SO_2(g) + NO_2(g) \rightleftharpoons SO_3(g) + NO(g)$

	$SO_2(g)$	$NO_2(g)$	$SO_3(g)$	$NO(g)$
amounts	0.400 mol	0.0500 mol	0.300 mol	0.200 mol
[ ]	0.400 mol/L	0.0500 mol/L	0.300 mol/L	0.200 mol/L

1.00L

$$K_{eq} = \frac{[SO_3][NO]}{[SO_2][NO_2]}$$

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

$$K_{eq} = 3.00$$

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7. $2NO_2 \rightleftharpoons N_2O_4$

$$K_{eq} = \frac{[N_2O_4]}{[NO_2]^2} \rightarrow [N_2O_4] = K_{eq}[NO_2]^2$$

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

$$= 1.15(0.50)^2$$

$$= 0.288$$

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8. $CO_2 + H_2 \rightleftharpoons CO + H_2O$

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

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

$$K_{eq} = 1.29$$

9. $2NH_3 \rightleftharpoons N_2 + 3H_2$ 1L

a)+b)

	$2NH_3$	N_2	$3H_2$
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) $K_{eq} = \frac{[N_2][H_2]^3}{[NH_3]^2}$

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

$K_{eq} = 0.00422$

d) S : ↑ $[H_2]$ H: use H_2
R : ↓ $[H_2]$ O: shift left

e) + f) only a change in temp will affect the value of Keq

10. $CO_2 + H_2 \rightleftharpoons H_2O + CO$ 1L

a)

	CO_2	H_2	H_2O	CO
	/	/	/	/
	0.5 mol	0.5 mol	∅	∅
	$0.5 - x$	$0.5 - x$	x	x
	$0.5 - x$	$0.5 - x$	x	x

Let x represent the amount of H_2O that forms

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

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

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

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

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

$K_{eq} = 2.00$

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

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

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

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

extraneous

$0.5 - x \not< 0$

$\therefore x \leq 0.5$

$\therefore [CO_2] = 0.5 - x$

$= 0.5 - 0.293$

$= 0.207$ mol/L

$\therefore [H_2] = 0.207$ mol/L

$[H_2O] = 0.293$ mol/L

$[CO] = 0.293$ mol/L

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

11. a) $H_2 + CO_2 \rightleftharpoons H_2O + CO$

	H_2	CO_2	H_2O	CO
amounts	1.17 mol	1.17 mol	1.33 mol	1.33 mol
[ ]s	0.117 mol	0.117 mol	0.133 mol	0.133 mol

10.0L

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

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

$$K_{eq} = 1.29$$

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

d) no effect, only temperature affects Keq

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

	H_2	CO_2	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

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

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

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

$$2.90 = 1.00 + x$$

$x = 1.9025$ mol H_2O injected