Quiz 15.3 – ICE Tables

Name: Key

Question 1

Consider the reaction: $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$ $K_C = 5.63 \times 10^{-4}$

Give the equilibrium concentrations of all species if excess C(s) is places in a chamber with $[H_2O(g)] = 0.250\ M$

	H_2O	CO	H_2
I	0.250 M	∅	∅
C	$-x$	$+x$	$+x$
E	$0.250-x$	x	x

$$5.63\cdot10^{-4} = \frac{x^2}{0.250 - x}$$ ← assume x is small

$0.250 \cdot 5.63\cdot10^{-4} = x^2 \rightarrow x = 0.0119$

It is 4.76% of 0.250 ✓

$[H_2O] = 0.238\ M$ $[CO] = 0.0119\ M$ $[H_2] = 0.0119\ M$

Question 2

Consider the reaction: $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ $K_C = 0.0160$

An amount of pure $PCl_5(g)$ is placed in an empty chamber. After equilibrium is reached, the product concentrations are measured as: $[PCl_3] = [Cl_2] = 0.0134\ M$
What are the initial and equilibrium concentrations of PCl_5?

	PCl_5	PCl_3	Cl_2
I	x	0	0
C	−0.0134	+0.0134	+0.0134
E	x − 0.0134	0.0134	0.0134

$$0.0160 = \frac{0.0134 \cdot 0.0134}{x - 0.0134}$$

$0.0160 \cdot x - 2.14\cdot10^{-4} = 1.796\cdot10^{-4}$

$x = 0.0246$

$[PCl_5]_0 = 0.0246\ M$

[PCl5](eq) = 0.0246M-.0134M = 0.0112M

Question 3

Consider the reaction: $Cl_2(g) + Br_2(g) \rightleftharpoons 2\,BrCl(g)$ $K_C = 7.20$

Find the equilibrium concentrations if a chamber is charged with 0.500 *mol* $Br_2(g)$ and 0.0500 *mol* BrCl and the reaction is allowed to reach equilibrium

	Cl_2	Br_2	$BrCl$
I	0	0.5	0.05
C	$+x$	$+x$	$-2x$
E	x	$0.5+x$	$0.05-2x$

$$7.2 = \frac{(0.05 - 2x)^2}{x \cdot (0.5 + x)}$$

↓ Assume x is small

$$7.2 = \frac{0.05^2}{x \cdot 0.5} \rightarrow x = 6.94\cdot10^{-4}$$

$2x$ is only 2.8% of 0.05!

$V \cdot [Cl_2] = 6.94\cdot10^{-4}$ moles

$V \cdot [Br_2] = 0.501$ moles

$V \cdot [BrCl] = 0.0486$ moles