Equilibrium III **Name: _____**

1. In the equilibrium system shown all substances are aqueous: $Ag^+ + 2\ NH_3 \leftrightarrow Ag(NH_3)_2^+$
When the two solutions first came into contact, the following concentrations were present:

$[Ag^+] = 0.60$ M **$[NH_3] = 0.40$ M**

After a time an equilibrium is established and it was found that : $[Ag(NH_3)_2^+]_{EQ} = 0.19$ M

Find the K_{eq}.

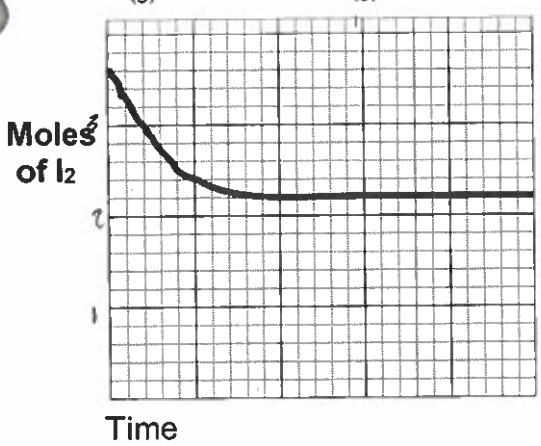
	Ag^+	$2NH_3$	$Ag(NH_3)_2^+$
I	0.60	0.40	0
C	-0.19	-0.38	+0.19
E	0.41	0.02	0.19

$$K_{eq} = \frac{[Ag(NH_3)_2^+]}{[Ag^+][NH_3]^2} = \frac{0.19}{(0.41)(0.02)^2}$$

$K_{eq} = 1158.5$ $\boxed{1 \times 10^3}$

2. The reaction below begun by mixing 3.6 moles of I_2 gas with 5.4 moles of Br_2 gas in a 1.0 litre container. The graph below shows the change in number of moles of I_2 over time as the reaction proceeded.

$I_{2\,(g)} + Br_{2\,(g)} \leftrightarrow 2\ IBr_{(g)}$ **Find the value of the K_{eq} for this reaction.**



	I_2	Br_2	$2IBr$
I	3.6M	5.4M	0
C	-1.4	-1.4	+2.8
E	2.2	4.0	2.8

$$K_{eq} = \frac{[IBr]^2}{[I_2][Br_2]} = \frac{(2.8)^2}{(2.2)(4.0)} = 0.89$$

3. Reactants A and B are mixed and their initial concentrations are [A] = 0.60M and [B] = 0.40 M. At equilibrium it is found that half of the reactant A was consumed. Calculated the K_{eq} for this system.

They react as follows: $A + B \leftrightarrow 2C + D$

	A	B	2C	D
I	0.60	0.40	0	0
C	-0.30	-0.30	+0.60	+0.30
E	0.30	0.10	0.60	0.30

$$K_{eq} = \frac{[C]^2[D]}{[A][B]} = \frac{(0.60)^2(0.30)}{(0.30)(0.10)}$$

$\boxed{K_{eq} = 3.6}$

Quadratic equation questions

4. At high temperatures $COBr_{2\,(g)}$ decomposes into $CO_{(g)}$ and $Br_{2(g)}$. In the lab you heat a 0.250M $COBr_{2\,(g)}$ sample until it decomposes. Find the equilibrium concentration for each reactant and product if the K_{eq} is 0.190 at this temperature.

	$COBr_2$	CO	Br_2
I	0.250	0	0
C	$-x$	$+x$	$+x$
E	$0.250-x$	x	x

$$K_{eq} = \frac{[CO][Br_2]}{[COBr_2]}$$

$$0.190 = \frac{x^2}{0.250-x}$$

$$0.0475 - 0.190x = x^2$$

$$0 = \underset{a}{x^2} + \underset{b}{0.190x} - \underset{c}{0.0475}$$

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

$$= \frac{-0.190 \pm \sqrt{(0.190)^2 - 4(1)(-0.0475)}}{2(1)}$$

$$= \frac{-0.190 \pm \sqrt{0.2261}}{2}$$

$x = -0.333$ impossible!

or

$x = 0.143$

$[COBr_2] = 0.250 - 0.143$
$= 0.107M$

$[CO] = 0.143M$

$[Br_2] = 0.143M$

5. Consider the reaction of H_2 reacting with I_2 to produce HI. Calculate the equilibrium concentrations of all 3 species if the initial concentration are; $K_{eq} = 54.3$

$[H_2]$ = 0.00623M
$[I_2]$ = 0.00414M
$[HI]$ = 0.0224M

$$H_2 + I_2 \rightarrow 2HI$$

I	0.00623	0.00414	0.0224
C	$-x$	$-x$	$+2x$
E	$0.00623-x$	$0.00414-x$	$0.0224+2x$

$[H_2] = 0.00467$
$[I_2] = 0.00258$
$[HI] = 0.0255$

$$K_{eq} = \frac{[HI]^2}{[I_2][H_2]}$$

$$54.3 = \frac{(0.0224+2x)^2}{(0.00623-x)(0.00414-x)}$$

$$54.3 = \frac{5.02\times10^{-4} + 0.0896x + 4x^2}{2.58\times10^{-5} - 0.0104x + x^2}$$

$$1.40\times10^{-3} - 0.56472x + 54.3x^2 = 5.02\times10^{-4} + 0.0896x + 4x^2$$

$$50.3x^2 - 0.654x + 8.98\times10^{-4} = 0$$

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

$$= \frac{0.654 \pm \sqrt{(-0.654)^2 - 4(50.3)(\quad)}}{2(50.3)}$$

8.98×10^{-4}

$x = \cancel{0.0114M}$ or $x = 0.00156M$