**Answers to Equilibrium Problems #2: Applying Le Chatelier's Principle**

1. $2\,HI_{(g)} \leftrightarrow H_{2(g)} + I_{2(g)}$

a) 0.97 mol of HI, 0.18 mol of $H_2$ and 0.12 mol of $I_2$ are **at equilibrium** in a 10.0 litre container. Calculate the value of Keq.

[HI] = n/V = 0.097 M  $[H_2]$ = 0.018 M  $[I_2]$ = 0.012 M

$$Keq = \frac{[H_2]\,[I_2]}{[HI]^2}$$

Keq = 0.023

b) 0.40 mol of hydrogen iodide gas are added to this system. In which direction will the equilibrium shift?

HI is a reactant, so the eq'm will shift to the right

c) Calculate the concentrations of the three chemicals when the system returns to equilibrium.

When 0.40 moles of HI are added, the number of ***moles*** of HI will be (0.97 + 0.40) = 1.37 moles. Before the equilibrium shifts to use up HI, the concentration of HI will be (C = n/V) = 0.137 M and the concentration of the other species will be the same as they were so fill in the ICE with these initial values:

| | $2\ HI_{(g)}$ $\Leftrightarrow$ | $H_{2\ (g)}$ + | $I_{2\ (g)}$ |
|---|---|---|---|
| **E'** | 0.97 | 0.018 | 00.12 |
| **I** | (0.97 + 0.40) | 0.018 | 00.12 |
| **C** | – 2 x | + x | + x |
| **E** | 0.137 - 2x | 0.018 + x | 0.012 + x |

$$Keq = \frac{[H_2]\,[I_2]}{[HI]^2}$$

Can we ignore the –2x?

$\frac{0.137}{0.023}$ is much less than 500, so we can't ignore –x, use quadratic

$$0.023 = \frac{[0.018 + x]\ [0.012 + x]}{[0.137 - 2x]^2}$$

$0.908\,x^2 + 0.0426x - 0.000216 = 0$

x = - 0.00462 (inadmissible) or x = 0.00463

so [HI] at eq'm = 0.128 mol/L = 0.13 M  $[H_2]$ = 0.0226 M = 0.023 M  $[I_2]$ = 0.0166 M = 0.017 M

2. Given the reaction: $CO_2$ (g) + $H_2$ (g) ↔ CO (g) + $H_2O$ (g)

At 900 ºC in a 10.0L flask, the concentration of each species is [CO] = 0.352 M, [$H_2O$] = 0.352 M, [$CO_2$] = 0.648 M and [$H_2$] = 0.148 M. If 4.00 moles of water vapour are added to the equilibrium system, what will the concentration of each species be when equilibrium is re-established?

**Step 1**: Calculate Keq at the original eq'm conditions using the concentrations given:

$$Keq = \frac{[CO]\,[H_2O]}{[H_2]\,[CO_2]}$$

Keq = 1.29

**Step 2**: When 4.00 moles of water vapour are added to the mixture, the eq'm will shift to left. The ***concentration*** of water vapour will be increased by (C = n/V) = 0.40 M

| | $H_2$ (g) | $CO_2$ (g) ⟺ | $H_2O$ (g) + | CO (g) |
|---|---|---|---|---|
| **E'** | 0.148 | 0.648 | 0.352 | 0.0.352 |
| **I** | 0.148 | 0.648 | (0.352 + 0.400) | 0.0.352 |
| **C** | + x | + x | - x | - x |
| **E** | 0.148 + x | 0.648 + x | 0.752 – x | 0.352 – x |

Can we ignore the – x?

$\frac{0.752}{1.29}$ is much less than 500, so we can't ignore –x, use quadratic

$$Keq = \frac{[CO]\,[H_2O]}{[H_2]\,[CO_2]}$$

$$1.29 = \frac{[0.352 - x]\,[0.752 - x]}{[0.148 + x]\,[0.648 + x]}$$

$0.29\,x^2 + 2.131x - 0.141 = 0$

x = 0.0655 or x = -7.41 (inadmissible)

so [CO] at eq'm = 0.2865 mol/L = 0.287 M

[$H_2O$] = 0.6865 M = 0.687 M

[$CO_2$] = 0.7135 M = 0.714 M

[$H_2$] = 0.2135 M = 0.214 M

3. $PCl_{5(g)} \leftrightarrow PCl_{3(g)} + Cl_{2(g)}$ K = 0.042

1.5 mol of $PCl_5$, 0.60 mol of $PCl_3$ and 0.60 mol of $Cl_2$ are put into a 1-litre flask. How many **moles** of each will be present at equilibrium?

**Step 1:** To find which way the reaction will proceed, calculate Q

$$Q = \frac{[PCl_3]\,[Cl_2]}{[PCl_5]}$$

Q = 0.24 Q > Keq, so we have more products than we "should", and the reaction will proceed to the left (⟵). See page 354 in your text if you are having trouble with this.

| | $PCl_{5\,(g)}$ ⟺ | $PCl_{3\,(g)}$ + | $Cl_{2\,(g)}$ |
|---|---|---|---|
| **I** | 1.5 | 0.60 | 0.60 |
| **C** | + x | – x | – x |
| **E** | 1.5 + x | 0.60 – x | 0.60 – x |

Can we ignore the + x?

$\frac{1.5}{0.042}$ is much less than 500, so we can't ignore x terms, use quadratic

$$Keq = \frac{[PCl_3]\,[Cl_2]}{[PCl_5]}$$

$$0.042 = \frac{[0.60 - x]\,[0.60 - x]}{[1.5 + x]}$$

$1\,x^2 - 1.242\,x + 0.30 = 0$

x = 0.328 or x = 0.9123 (inadmissible because $[PCl_3]$ will be (0.60 – 0.9123), which is negative)

so $[PCl_5]$ at eq'm = 1.83 M $[PCl_3]$ = 0.272 M $[Cl_2]$ = 0.272 M

4. $CO_{(g)}$ + $Cl_{2(g)}$ $\leftrightarrow$ $COCl_{2(g)}$

1.0 mol of each of CO, $Cl_2$ and $COCl_2$ are present at equilibrium in a 5-litre container. How many **moles** of each will be present if the volume of the container is decreased to 3.00 litres and the system is allowed to reach equilibrium again?

1.0 mol of CO, $Cl_2$ and $COCl_2$ are **at equilibrium** in a 5.0 litre container. Calculate the value of Keq.

[CO] = n/V = 0.20 M    $[Cl_2]$ = 0.20 M    $[COCl_2]$ = 0.20 M

$$Keq = \frac{[COCl_2]}{[CO]\ [Cl_2]}$$

Keq = 5.0

When the volume of the container is decreased, the pressure will increase so the eq'm will shift to the right to decrease the number of particles of gas. The new concentrations of the gases will be (C = n/V 1.0 mol/3.0 L = 0.333 M). Complete an E'ICE table to calculate the new eq'm concentrations.

| | CO (g) + | Cl₂ (g) ⟺ | COCl₂ (g) |
|---|---|---|---|
| **E'** | 0.20 | 0.20 | 0.20 |
| **I** | 0.333 | 0.333 | 0.333 |
| **C** | – x | – x | + x |
| **E** | 0.333 – x | 0.333 – x | 0.333 + x |

$$Keq = \frac{[COCl_2]}{[CO]\ [Cl_2]}$$

$$5.0 = \frac{[0.333 + x]}{[0.333 - x]\ [0.333 - x]}$$

Can we ignore the – x?

$\frac{0.333}{5.0}$ is much less than 500, so we can't ignore –x, use quadratic

$5.0\,x^2 - 4.33\,x + 0.221 = 0$

x = 0.054 or x = 0.812 (inadmissible because $[Cl_2]$ would be negative)

so [CO] at eq'm = 0.28 mol/L    $[Cl_2]$ = 0.28 M    $[COCl_2]$ = 0.39

Now, convert these concentrations to number of moles of each gas by multiplying by 3.0 L (n = C x V)

mol CO = 0.84 mol    mol $Cl_2$ = 0.84 mol    mole $COCl_2$ = 1.2 mol    (2 sig digs)

5. This is a challenging problem, just for fun. Be sure to keep track of the mole ratios and changes in concentration.

$2\ NO_{(g)} + O_{2(g)} \leftrightarrow 2\ NO_{2(g)}$ Keq = 403.7

An equilibrium mixture in a 2.00 L container is composed of 0.24 mol of NO, 0.086 mol of $O_2$ and 1.00 mol of $NO_2$.
How many moles of oxygen would have to be added to the mixture to increase the amount of $NO_2$ to 1.20 moles when equilibrium is re-established?

| | 2 NO (g) + | $O_2$ (g) → | 2 $NO_2$ (g) |
|---|---|---|---|
| **E'** | 0.12 | 0.043 | 0.50 |
| **I** | 0.12 | 0.043 + x | 0.50 |
| **C** | – 0.10 | – 0.050 | + 0.10 |
| **E** | 0.020 | ( – 0.007 + x ) | 0.60 |

**Logic:**

- we want to increase the number of moles of $NO_2$ to 1.20 moles
- because the container is a 2.0 L container, the ***concentration*** will increase by (C = n/V) 0.60 M
- so, the [$NO_2$] will ***change*** by (0.60 – 0.50) = 0.10 mol/L, so this is our change term
- because of mole ratios (both NO and $NO_2$ have a mole ratio of 2), NO will also change by 0.10 M, but $O_2$ has a mole ratio of only 1, so its change will be half as much, or 0.050 M

Finally, substitute the final equilibrium concentrations into the Keq expression and solve for x:

$$Keq = \frac{[NO_2]^2}{[NO]^2\ [O_2]}$$

$$403.7 = \frac{[0.60]^2}{[0.020]^2\ [-0.007 + x]}$$

Expand and solve for x = 2.24

Therefore, the concentration change is 2.24 M

Then n = C V

= 2.24 M x 2.0 L

= 4.48 mole

So, 4.48 moles of $O_2$ must be added to bring the final amount of $NO_2$ to 1.20 moles