Calculating Concentration at Equilibrium

(TICE)

P 319 # 18,19,20,21,14,15,17 (Do in this order)

18. $C = \frac{n}{V}$

~~$[HBr]_i = \frac{0.090 \text{ mol}}{2 \text{ L}} = 0.045 \text{ mol/L}$~~

~~*Answer:* The equilibrium constant is so low that the concentration at equilibrium will be the same as the initial concentration, that is, 0.045 mol/L for the hydrogen bromide (HBr) and 0 mol/L for the products.~~

19. *1. Calculation of the initial concentrations:*

$C = \frac{n}{V}$

$[C]_i = [D]_i = \frac{1.0 \text{ mol}}{2 \text{ L}} = 0.50 \text{ mol/L}$ | $[E] = \frac{0.1}{2} = 0.05 \text{ mol/L}$

✓

2. Recording of data and use of the ICE table:

	Concentration (mol/L)	$C_{(g)}$ +	$D_{(g)}$ ⇌	$E_{(g)}$ +	$2F_{(g)}$
T					
I	Initial (C_i)	0.50	0.50	0	0
C	Variation (ΔC)	-0.05	-0.05	+0.05	+(2)0.05
E	Equilibrium (C_{eq})	0.45	0.45	0.05	0.1

3. Calculation of the equilibrium constant:

$$K_c = \frac{[C]^c \cdot [D]^d}{[A]^a \cdot [B]^b} = \frac{[E] \cdot [F]^2}{[C] \cdot [D]} = \frac{(0.05) * (0.1)^2}{(0.45) * (0.45)}$$

$$= 0.025$$

Answer: The value of the equilibrium constant is 0.0025

20. *1. Calculation of the initial concentrations:*

$$C = \frac{n}{V}$$

$$[H_2]_i = \frac{5.00 \times 10^{-3}\ \text{mol}}{5\ \text{L}} = 0.001\ 00\ \text{mol/L}$$

$$[I_2]_i = \frac{1.00 \times 10^{-2}\ \text{mol}}{5\ \text{L}} = 0.002\ 00\ \text{mol/L}$$

2. Recording of data and use of the ICE table:

	Concentration (mol/L)	$H_{2\,(g)}$ +	$I_{2\,(g)}$ ⇌	$2\,HI_{(g)}$
T				
I	Initial (C_i)	0.001 00	0.002 00	0
C	Variation (ΔC)	−0.000 935	−0.000 935	+2(0.000 935)
E	Equilibrium (C_{eq})	0.000 065	0.001 065	0.001 87

3. Calculation of the equilibrium constant:

$$K_c = \frac{[C]^c \cdot [D]^d}{[A]^a \cdot [B]^b} = \frac{[HI]^2}{[H_2] \cdot [I_2]}$$

$$\frac{0.001\ 87^2}{0.000\ 065 \cdot 0.001\ 065} = 50.5$$

Answer: The value of the equilibrium constant at 448°C is 50.5.

21. 1. *Calculation of the initial concentrations:*

$$C = \frac{n}{V}$$

$$[NO]_i = [O_3]_i = \frac{4.6 \text{ mol}}{2 \text{ L}} = 2.3 \text{ mol/L}$$

$$[NO_2]_i = \frac{2.4 \text{ mol}}{2 \text{ L}} = 1.2 \text{ mol/L}$$

2. *Recording of data and use of the ICE table:*

	Concentration (mol/L)	$NO_{(g)}$ +	$O_{3(g)}$ ⇌	$NO_{2(g)}$ +	$O_{2(g)}$
I	Initial (C_i)	2.3	2.3	0	0
C	Variation (ΔC)	−1.2	−1.2	+1.2	+1.2
E	Equilibrium (C_{eq})	1.1	1.1	1.2	1.2

T

3. *Calculation of the equilibrium constant:* ✓

$$K_c = \frac{[C]^c \cdot [D]^d}{[A]^a \cdot [B]^b} = \frac{[NO_2] \cdot [O_2]}{[NO] \cdot [O_3]} = \frac{1.2 \cdot 1.2}{1.1 \cdot 1.1} = 1.19$$

Answer: The value of the equilibrium constant is 1.2.

14. *1. Calculation of initial concentrations:*

$C = \frac{n}{V}$

$[I_2]_i = [Cl_2]_i = \frac{0.83 \text{ mol}}{10 \text{ L}} = 0.083 \text{ mol/L}$

2. Recording of data and use of the ICE table:

Concentration (mol/L)	$I_{2\,(g)}$ +	$Cl_{2\,(g)}$ ⇌	$2\,ICl_{(g)}$
T I C E			
Initial (C_i)	0.083	0.083	0
Variation (ΔC)	$-x$	$-x$	$+2x$
Equilibrium (C_{eq})	$0.083 - x$	$0.083 - x$	$2x$

3. Calculation of concentrations of each substance at equilibrium:

$$K_c = \frac{[C]^c \cdot [D]^d}{[A]^a \cdot [B]^b} = \frac{[ICl]^2}{[I_2] \cdot [Cl_2]}$$

$$82 = \frac{(2x)^2}{(0.083 - x) \cdot (0.083 - x)}$$

This second-degree equation is of the type $ax^2 + bx + c = 0$. Therefore, it must be rewritten as a quadratic equation.

$82(0.006\,889 - 0.166x + x^2) = (2x)^2$

$0.564\,898 - 13.612x + 82x^2 = 4x^2$

$78x^2 - 13.612x + 0.564\,898 = 0$

To find the values of x, we can use:

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

$$= \frac{-(-13.612) \pm \sqrt{-13.612^2 - 4(78 \cdot 0.567\,898)}}{2 \cdot 78}$$

$= 0.1055$ or 0.069 (the 2nd value is the correct value)

✓

4. Calculation of the concentrations of each of the substances at equilibrium:

$[I_2]_{eq} = [Cl_2]_{eq} = 0.083 - (0.069) = \boxed{0.014 \text{ mol/L}}$

$[ICl]_{eq} = 2(0.069) = \boxed{0.138 \text{ mol/L}}$

Answer: At equilibrium, the concentration of the iodine (I_2) is 0.014 mol/L, that of chlorine (Cl_2) is 0.014 mol/L and that of iodine chloride (ICl) is 0.138 mol/L.

15. *1. Recording of data and use of the ICE table:*

T	Concentration (mol/L)	$2HF_{(g)}$ ⇌	$H_{2(g)}$ +	$F_{2(g)}$
I	Initial (C_i)	x	0	0
C	Variation (ΔC)	$-2(0.045)$	$+0.045$	$+0.045$
E	Equilibrium (C_{eq})	$x - 0.090$	0.045	0.045

2. Calculation of the concentrations of each substance at equilibrium:

$$K_c = \frac{[C]^c \cdot [D]^d}{[A]^a \cdot [B]^b} = \frac{[H_2] \cdot [F_2]}{[HF]^2}$$

$$4 = \frac{(0.045) \cdot (0.045)}{(x - 0.090)^2}$$

This second-degree equation is of the type $ax^2 + bx + c = 0$. Therefore, it must be rewritten as a quadratic equation.

$$4(x^2 - 0.18x + 0.0081) = 0.002\,025$$
$$4x^2 - 0.72x + 0.0324 = 0.002\,025$$
$$4x^2 - 0.72x + 0.030\,375 = 0$$

To find the values of x, we can use:

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

$$= \frac{-(-0.72) \pm \sqrt{-0.72^2 - 4(4 \cdot 0.030\,375)}}{2 \cdot 4}$$

$= 0.1125$ or 0.0675 (the 1st value is the correct value)

Therefore, the initial concentration is 0.1125 mol/L. $= x$

Answer: There was 0.11 mol of hydrogen fluoride (HF) in the reaction jar at the beginning of the reaction.

✓

17. *1. Calculation of the initial concentrations:*

$$C = \frac{n}{V}$$

$$[CO]_i = \frac{0.055 \text{ mol}}{5 \text{ L}} = 0.011 \text{ mol/L}$$

$$[Cl_2]_i = \frac{0.072 \text{ mol}}{5 \text{ L}} = 0.0144 \text{ mol/L}$$

2. Recording of data and use of the ICE table:

	Concentration (mol/L)	$CO_{(g)}$ +	$Cl_{2(g)}$ ⇌	$COCl_{2(g)}$
T				
I	Initial (C_i)	0.011	0.0144	0
C	Variation (ΔC)	$-x$	$-x$	$+x$
E	Equilibrium (C_{eq})	$0.011 - x$	$0.0144 - x$	x

3. Calculation of the concentrations of each substance at equilibrium:

$$K_c = \frac{[C]^c \cdot [D]^d}{[A]^a \cdot [B]^b} = \frac{[COCl_2]}{[CO] \cdot [Cl_2]}$$

$$0.20 = \frac{x}{(0.011 - x) \cdot (0.0144 - x)}$$

This second-degree equation is of the type $ax^2 + bx + c = 0$. Therefore, it must be rewritten as a quadratic equation.

$0.20(0.000\,158\,4 - 0.0254x + x^2) = x$
$0.000\,031\,68 - 0.005\,08x + 0.20x^2 = x$
$0.20x^2 - 1.005\,08x + 0.000\,031\,68 = 0$
To find the values of x, we can use:

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

$$= \frac{-(-1.005\,08) \pm \sqrt{-1.005\,08^2 - 4(0.20 \cdot 0.000\,031\,68)}}{2 \cdot 0.20}$$

$= 5.025$ or $0.000\,031\,52$ (the 2nd value is the correct value)

4. Calculation of the concentrations of each substance at equilibrium:

$[CO]_{eq} = 0.011 - x = 0.010\,97$ mol/L
$[Cl_2]_{eq} = 0.0144 - x = 0.014\,37$ mol/L
$[COCl_2]_{eq} = 3.2 \times 10^{-5}$ mol/L

Answer: At equilibrium, the concentration of the carbon monoxide (CO) is 0.011 mol/L, that of chlorine (Cl_2) is 0.0144 mol/L and that of phosgene ($COCl_2$) is 3.2×10^{-5} mol/L.