

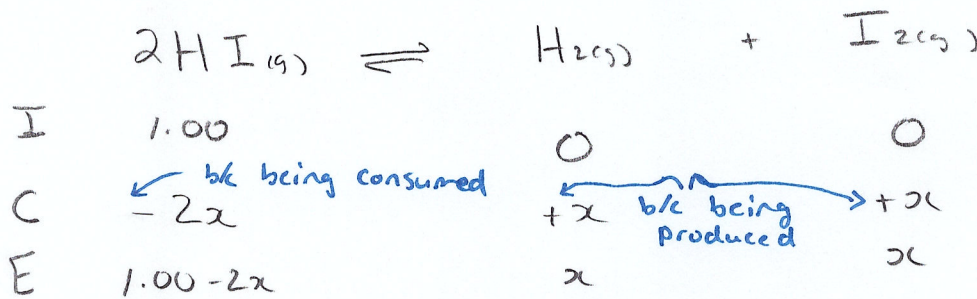
ICE Tables

- If we place only reactants in a container, the system will reach an equilibrium
 - ✗ ○ Therefore, we need to be able to calculate the equilibrium concentrations based on the initial concentrations
- ICE tables are a convenient way to organize and calculate the changes in the concentrations of products and reactants for a system that reaches equilibrium
 - "ICE" stands for the initial concentration, change in concentration, and equilibrium concentration of all chemicals involved in a reaction system arrange into a table format
 - ✗ ○ ICE table only contain concentration values

EXAMPLES:

1. Consider the reaction for the decomposition of hydrogen iodide at 448°C. The initial concentration of $\text{HI}_{(g)}$ was 1.00 mmol/L. Once an equilibrium was established, the concentration of $\text{HI}_{(g)}$ was measured to be 0.078 mmol/L. Calculate the equilibrium constant (K_c).

→ need equilibrium concentrations ∴ ICE table!



← b/c no products initially present

Set table up with variables first!

* table set up in mmol/L *

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{(x)(x)}{(1.00 - 2x)^2} = \frac{x^2}{(1.00 - 2x)^2}$$

but $[\text{HI}]_{\text{equ.}} = 1.00 - 2x = 0.078$

$$1.00 - 0.078 = 2x$$

$$0.922 = 2x$$

$$0.461 = x$$

$$K_c = \frac{(0.461)^2}{(0.078)^2} = 34.9311 \dots$$

$K_c = 35$

2. Nitrogen dioxide can break down into nitrogen monoxide and oxygen. The equilibrium constant for this reaction is $K_c = 0.40$. If the equilibrium concentration of $\text{NO}_{2(g)}$ is 0.20 mol/L and the equilibrium concentration of $\text{NO}_{(g)}$ is 1.0 mol/L , what is the equilibrium concentration of $\text{O}_{2(g)}$? $\rightarrow \therefore$ don't use ICE table!

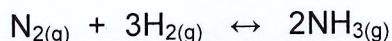


$$\therefore K_c = \frac{[\text{O}_2][\text{NO}]^2}{[\text{NO}_2]^2} \Rightarrow [\text{O}_2] = \frac{K_c [\text{NO}_2]^2}{[\text{NO}]^2}$$

$$[\text{O}_2] = \frac{(0.40)(0.20)^2}{(1.0)^2} = 0.016$$

$$\boxed{[\text{O}_2] = 0.016 \text{ mol/L}}$$

3. Ammonia gas ($\text{NH}_3(\text{g})$) is produced in industrial amounts for use in the production of fertilizers. The most well-known industrial process to produce ammonia gas in large quantities is using the Haber-Bosch process as shown below.



The equilibrium constant for the Haber-Bosch process is 0.469. If 20.0 mol of $\text{N}_2(\text{g})$ and 75.0 mol of $\text{H}_2(\text{g})$ are placed in a 5.00L reaction vessel and 33.0 mol of ammonia is present at equilibrium, what is the equilibrium concentration of hydrogen gas?

* ICE table needs to be set up with concentration values!

$$C = \frac{n}{V} \quad \therefore [\text{N}_2] = \frac{20.0 \text{ mol}}{5.0 \text{ L}} = 4.00 \text{ mol/L}$$

$$[\text{H}_2] = \frac{75.0 \text{ mol}}{5.0 \text{ L}} = 15.0 \text{ mol/L}$$

	$\text{N}_2(\text{g})$	+	$3\text{H}_2(\text{g})$	\rightleftharpoons	$2\text{NH}_3(\text{g})$
I	4.00		15.0		0
C	-x		-3x		+2x
E	4.00-x		15.0-3x		2x

but $[\text{NH}_3]_{\text{equil.}} = \frac{33.0 \text{ mol}}{5.0 \text{ L}} = 6.6 \text{ mol/L} = 2x$ ← from ICE table

$\therefore x = 3.3 \text{ mol/L}$

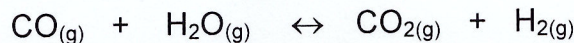
$$[\text{H}_2]_{\text{equil.}} = \underbrace{15.0 - 3x}_{\substack{\uparrow \\ \text{from ICE} \\ \text{table}}} = 15.0 - 3(3.3) = 5.1 \text{ mol/L}$$

Note: we didn't even need the $K_c = 0.469$!

$[\text{H}_2]_{\text{equil.}} = 5.1 \text{ mol/L}$

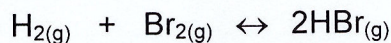
Practice Problems

1. Consider the equilibrium system:



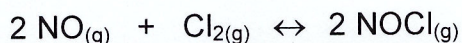
Initially, 0.75 mol of water and 0.60 mol of carbon monoxide are placed in a 3.0L reaction vessel. At equilibrium, 0.30 mol of carbon dioxide is present. Calculate the equilibrium constant for the system. **[0.67]**

2. Consider the equilibrium system:



Initially, 0.25 mol of hydrogen and 0.25 mol of bromine are placed in a 500mL reaction vessel. The equilibrium constant for the reaction is 50. Calculate the equilibrium concentration of bromine when the equilibrium concentration of hydrogen bromide is 0.78mol/L. **[0.11mol/L]**

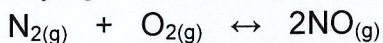
3. Consider the equilibrium system:



Initially, 5.0 mol of $\text{NO}_{(g)}$ and 5.0 mol of $\text{Cl}_{2(g)}$ were added to a 2.0 L container. As a result of the reaction, the equilibrium concentration of $\text{NOCl}_{(g)}$ became 0.96 mol/L. Determine the value of the equilibrium constant (K_c) for this reaction. **[0.19]**

End of Practice Problems

4. Nitrogen oxides from exhaust gases are a serious pollution problem. An environmental chemist is studying the following equilibrium reaction:



At a temperature of the exhaust gases from a particular engine, the value of K_c is 4.2×10^{-8} . The chemist puts 0.17 mol of nitrogen and 0.076 mol of oxygen in a rigid 2.0L cylinder. What is the concentration of nitrogen monoxide in the mixture at equilibrium?

← initial amounts
∴ use ICE table!

* First need concentration values!

$$C = \frac{n}{V} \quad \therefore C_{\text{N}_2} = \frac{0.17 \text{ mol}}{2.0 \text{ L}} = 0.085 \text{ mol/L}$$

$$C_{\text{O}_2} = \frac{0.076 \text{ mol}}{2.0 \text{ L}} = 0.038 \text{ mol/L}$$

	$\text{N}_{2(g)}$	+	$\text{O}_{2(g)}$	\rightleftharpoons	$2\text{NO}_{(g)}$
I	0.085		0.038		0
C	$0.085 - x$		$0.038 - x$		$+ 2x$
E	$0.085 - x$		$0.038 - x$		$2x$

$$K_c = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} = \frac{(2x)^2}{(0.085-x)(0.038-x)}$$

↘ "foiled" / expanded

$$4.2 \times 10^{-8} = \frac{4x^2}{0.00323 - 0.123x + x^2}$$

a quadratic equation!

- Some equilibrium calculations can become very complex and time consuming to solve
- The previous example resulted in a quadratic equation which can be solved using the quadratic equation or by graphing
- However, for this example and ones that are more complex than a quadratic equation, we can use the **approximation rule**
- * • The **approximation rule** states that if $K_c \times 1000$ is less than the initial concentration of all reactants, then we can assume that the initial concentration of the reactants is the equilibrium concentration as well because K_c is so small which means hardly any reactants are being consumed
- Now let's finish our last example...

- first check to see if the approximation rule can apply

$$4.2 \times 10^{-8} \times 1000 = 4.2 \times 10^{-5} < 0.085 \text{ mol/L} \quad \text{and} \quad 0.038 \text{ mol/L}$$

\therefore use approximation rule!

$$4.2 \times 10^{-8} = \frac{4x^2}{(0.085-x)(0.038-x)}$$

can assume

$$0.085 - x \approx 0.085$$

$$0.038 - x \approx 0.038$$

$$4.2 \times 10^{-8} = \frac{4x^2}{(0.085)(0.038)}$$

$$4.2 \times 10^{-8} = \frac{4x^2}{0.00323} \Rightarrow 1.3566 \times 10^{-10} = 4x^2$$

$$3.3915 \times 10^{-11} = x^2$$

$$5.9236 \dots \times 10^{-6} = x$$

$$[\text{NO}]_{\text{equ.}} = 2x = 2(5.9236 \dots \times 10^{-6} \text{ mol/L})$$

$$[\text{NO}]_{\text{equ.}} = 1.2 \times 10^{-5} \text{ mol/L}$$

Now try pg. 658 #24 & pg. 660# 25-29