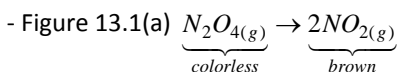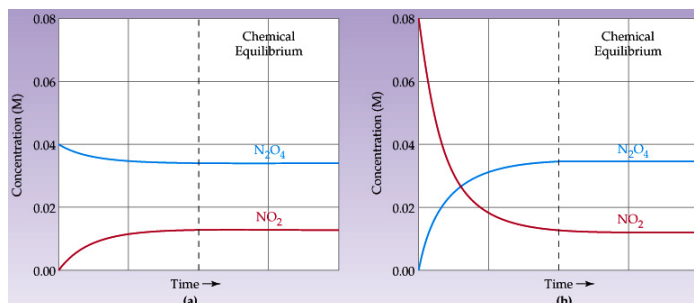# Chapter 15 – Chemical Equilibrium

- **15.1 The Concept of Equilibrium**

**Figure: 13.1** from *Chemistry* by McMurray & Fey



- Figure 13.1(a) $\underbrace{N_2O_{4(g)}}_{colorless} \rightarrow \underbrace{2NO_{2(g)}}_{brown}$
  - -- we start with reactant, $N_2O_4$, so the solution is colorless
  - -- as time progresses we generate product, $NO_2$
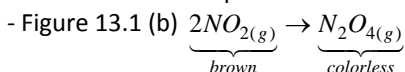  - -- after a certain length of time:
    - --- the concentration of $N_2O_4$ stops decreasing
    - --- the concentration of $NO_2$ stops increasing
    - --- the solution goes back and forth between clear and brown
    - --- at this point the forward rate = reverse rate
    - --- the production of $NO_2$ and the consumption of $N_2O_4$ is equal
- Figure 13.1 (b) $\underbrace{2NO_{2(g)}}_{brown} \rightarrow \underbrace{N_2O_{4(g)}}_{colorless}$
  - -- this time we start with $NO_2$
  - -- as time progresses we once again reach a point in which:
    - --- the concentration of $NO_2$ stops decreasing
    - --- the concentration of $N_2O_4$ stops increasing
    - --- the solution goes back and forth between brown and clear
    - --- at this point the forward rate = reverse rate
    - --- the production of $N_2O_4$ and the consumption of $NO_2$ is equal
- this is an example of <u>chemical equilibrium</u> or dynamic equilibrium: a phenomenon in which the concentrations of reactants and products remain constant over time
- NOTE: this does not mean that the concentrations go to zero
- we can rewrite the above eqn as $N_2O_{4(g)} \, 2NO_{2(g)}$

- **15.2 The Equilibrium Constant**
  - Evaluating $K_c$

$$aA + bB \rightleftharpoons cC + dD \qquad K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

  - -- this is called the equilibrium constant expression or mass action expression
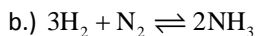  - -- we call $K_c$ the equilibrium constant since there is a certain set of concentrations for the species in our general equation which corresponds to equilibrium
  - -- law of mass action: the chemical equilibrium expression will give rise to a characteristic value for a given temperature

- Example: Write the equilibrium expression for the forward and the reverse reactions for the following:

a.) $4A + 3B_2 \rightleftharpoons 2C + D$
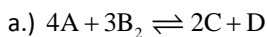
$$K_c = \frac{[C]^2[D]}{[A]^4[B]^3} \quad K_c' = \frac{[A]^4[B]^3}{[C]^2[D]}$$

b.) $3H_2 + N_2 \rightleftharpoons 2NH_3$

$$K_c = \frac{[NH_3]^2}{[H_2]^3[N_2]} \quad K_c' = \frac{[H_2]^3[N_2]}{[NH_3]^2}$$

- In both examples the equilibrium constant in the forward direction is the inverse of

the equilibrium constant in the reverse direction: $K_c = \dfrac{1}{K_c'}$

-- this means that when we reverse an eq reaction we must take the inverse of it's eq constant to get the reverse eq constant

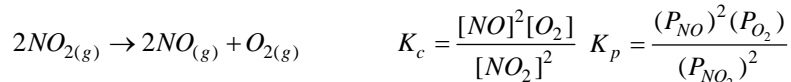- Equilibrium Constants in Terms of $K_p$

- if A, B, C, and D are gases then we would write: $K_p = \dfrac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$

-- Example: What is $K_c$ and $K_p$ for the following reaction?

$$2NO_{(g)} + Cl_{2(g)} \rightleftharpoons 2NOCl_{(g)}$$
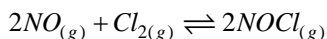
$$K_c = \frac{[NOCl]^2}{[NO]^2[Cl_2]} \qquad K_p = \frac{(P_{NOCl})^2}{(P_{NO})^2(P_{Cl_2})}$$

- How are $K_c$ and $K_p$ related?

$$2NO_{2(g)} \rightarrow 2NO_{(g)} + O_{2(g)} \qquad K_c = \frac{[NO]^2[O_2]}{[NO_2]^2} \quad K_p = \frac{(P_{NO})^2(P_{O_2})}{(P_{NO_2})^2}$$

-- looking at the ideal gas law: $\qquad P = \dfrac{n}{V}RT = \underline{M}RT \rightarrow \underline{M} = P\dfrac{1}{RT}$

-- plugging this expression into $K_c$ gives:

$$K_c = \frac{[NO]^2[O_2]}{[NO_2]^2} = \frac{\left(\dfrac{P_{NO}}{RT}\right)^2\left(\dfrac{P_{O_2}}{RT}\right)}{\left(\dfrac{P_{NO_2}}{RT}\right)^2} = \frac{(P_{NO})^2(P_{O_2})}{(P_{NO_2})^2}\frac{1}{RT}$$

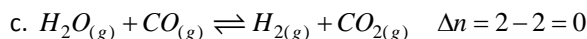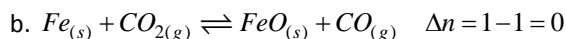-- therefore the relationship btwn $K_c$ & $K_p$ for this equation is:

$$K_c = K_p\frac{1}{RT} \rightarrow K_p = K_c(RT)$$

-- In general, $K_p = K_c(RT)^{\Delta n}$ where $\Delta n$ = moles$_{(g)products}$ - moles$_{(g)reactants}$

--- in this example $\Delta n = 1$

--- when $\Delta n = 0$ then $K_c = K_p$

-- Example: For which of the following reaction will $K_c = K_p$?

a. $2SO_{2(g)} + O_{(g)} \rightleftharpoons 2SO_{3(g)} \quad \Delta n = 2-3 = -1$

b. $Fe_{(s)} + CO_{2(g)} \rightleftharpoons FeO_{(s)} + CO_{(g)} \quad \Delta n = 1-1 = 0$

c. $H_2O_{(g)} + CO_{(g)} \rightleftharpoons H_{2(g)} + CO_{2(g)} \quad \Delta n = 2-2 = 0$
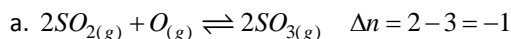
therefore for b. & c.

-- Example: What is the $K_p$ of the reaction below at 325°C given $K_c$ = 5.0?

$$Cl_{2(g)} + CO_{(g)} \rightleftharpoons COCl_{2(g)}$$
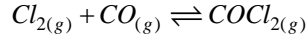
$$K_p = K_c(RT)^{\Delta n} = 5.0\,\tfrac{L}{mol}\left(0.08206\,\tfrac{L \cdot atm}{mol \cdot K} \times (325 + 273.15)K\right)^{(1-2)}$$

$$K_p = 0.10$$

<5>

## • 15.3 Understanding & Working with Equilibrium Constants

- The Magnitude of Equilibrium Constants
  - -- $K_c > 10^3$, products predominate over reactants:
    - --- reaction proceeds to completion - generating product
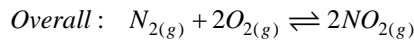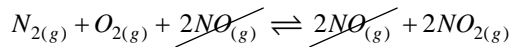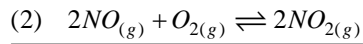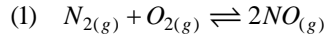    - --- very little reactant is left
  - -- $K_c < 10^{-3}$, reactants predominate over products: $K_c = \dfrac{[\text{products}]}{[\text{reactants}]}$
    - --- reaction proceeds hardly at all
    - --- very little product is generated
  - -- $10^3 > K_c > 10^{-3}$, both reactants and products are present at eq

### - Handling Combined Equations

$$(1) \quad N_{2(g)} + O_{2(g)} \rightleftharpoons 2NO_{(g)}$$

$$(2) \quad 2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$$

$$N_{2(g)} + O_{2(g)} + 2\cancel{NO}_{(g)} \rightleftharpoons 2\cancel{NO}_{(g)} + 2NO_{2(g)}$$

$$Overall: \quad N_{2(g)} + 2O_{2(g)} \rightleftharpoons 2NO_{2(g)}$$

- the mass action exp'n for the overall reaction is: $K_C = \dfrac{[NO_2]^2}{[N_2][O_2]}$

- the mass action exp'n for (1) is: $K_1 = \dfrac{[NO]^2}{[N_2][O_2]}$

- the mass action exp'n for (2) is: $K_2 = \dfrac{[NO_2]^2}{[NO]^2[O_2]}$

- we derive the mass action exp'n from (1) & (2) for the overall reaction thru $K_1$ x $K_2$

$$K_1 \times K_2 = \frac{\cancel{[NO]^2}}{[N_2][O_2]} \times \frac{[NO_2]^2}{\cancel{[NO]^2}[O_2]} = \frac{[NO_2]^2}{[N_2][O_2]^2}$$

- In general: $K_{overall}$ = $K_1$ x $K_2$ x ...
- Example: Calculate the eq const for $2D \rightleftharpoons A + 2B$ given the info below.

$$A + 2B \rightleftharpoons C \qquad K_c = 3.3$$

$$C \rightleftharpoons 2D \qquad K_c = 0.041$$

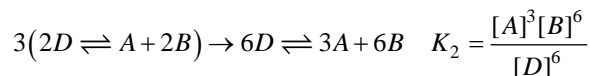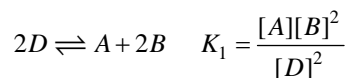- we need to look from the reverse direction for both of these reactions

$$C \rightleftharpoons A + 2B \qquad K_c = \frac{1}{3.3} = 0.303$$

$$2D \rightleftharpoons C \qquad K_c = \frac{1}{0.041} = 24.4$$

$$2D \rightleftharpoons A + 2B \qquad K_c = 0.303 \times 24.4 = 7.39$$

### - How does multiplying a reaction by a constant affect the eq. const?
- the eq. constant is raised by the exponential value of the constant

$$2D \rightleftharpoons A + 2B \quad K_1 = \frac{[A][B]^2}{[D]^2}$$

$$3(2D \rightleftharpoons A + 2B) \rightarrow 6D \rightleftharpoons 3A + 6B \quad K_2 = \frac{[A]^3[B]^6}{[D]^6}$$

$$\therefore K_2 = (K_1)^3$$

- Example: If the eq. constant at a given temperature is 2.4 x 10$^{-3}$ for

   $2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$ what is the eq. constant for the reactions below?

   a.) $SO_{2(g)} + \frac{1}{2}O_{2(g)} \rightleftharpoons SO_{3(g)}$ $\qquad\qquad$ $K_c = (2.4 \text{ x } 10^{-3})^{\frac{1}{2}} = 0.049$

   b.) $2SO_{3(g)} \rightleftharpoons 2SO_{2(g)} + O_{2(g)}$ $\qquad\qquad$ $K_c = (2.4 \text{ x } 10^{-3})^{-1} = 417$

   c.) $SO_{3(g)} \rightleftharpoons SO_{2(g)} + \frac{1}{2}O_{2(g)}$ $\qquad\qquad$ $K_c = (2.4 \text{ x } 10^{-3})^{-\frac{1}{2}} = 20.4$

   - *Summary of manipulating the mass action expression*

| Action | New Eq Constant |
|---|---|
| Reversing an equilibrium reaction | $K_c^{'} = \frac{1}{K_c}$ |
| Multiplying an equilibrium reaction by n | $K_c^{'} = (K_c)^n$ |
| Adding several eq reactions together | $K_c^{'} = K_1 \times K_2 \dots$ |

- **15.4 Heterogeneous Equilibria**
   - so far we have dealing with reactions in which everything is either in solution or is a gas - homogeneous eq
   - heterogeneous equilibrium: when the reactants/products are in more than one phase
   - the eq constant is independent of solid species - solids have no pressure and the concentration is constant
      (they are not in solution)

      -- Ex: $H_2O_{(g)} + C_{(s)} \rightleftharpoons CO_{(g)} + H_{2(g)}$ $\quad K_c = \frac{[CO][H_2]}{[H_2O]}$
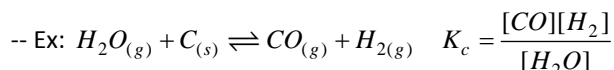
   - the eq constant is independent of $H_2O_{(l)}$ since it is the solvent and therefore in excess

      -- Ex: $CO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)}$ $\quad K_c = \frac{[H_2CO_3]}{[CO_2]}$
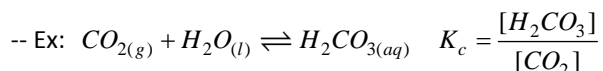
- **15.5 Calculating Equilibrium Constants**
   - *When Eq concentrations are known:* it is a matter of plugging them into the equilibrium expression
      -- Ex: What is the Kc for $2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$ if their equilibrium concentrations are [SO$_2$]=0.15 M,

         [O$_2$]=0.68, [SO$_3$]=1.5

      1. Write down the equilibrium expression symbolically:

         $$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

      2. Plug the give equilibrium values into the expression and solve:

         $$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{[1.5]^2}{[0.15]^2[0.68]} = 1.5 \times 10^2$$

   - *When initial concentrations and some equilibrium concentrations are known:* we use an "ICE" table,
      "I"nitial, "C"hange, "E"q & stoichiometry

-- Ex: What is the Kc for $2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$ if the initial concentrations of reactants were [SO$_2$]=0.150M and [O$_2$]=0.680M and the equilibrium concentration of the product is [SO$_3$]=0.050M?

1. Setup the ICE table

|  | SO$_2$ | O$_2$ | SO$_3$ |
|---|---|---|---|
| Initial | 0.15 | 0.68 | 0.0 |
| Change | -2x | -x | +2x |
| Eq | ? | ? | 0.05 |

2. Use the given equilibrium concentration to identify x

$2x = 0.050 \rightarrow x = 0.025$

3. Plug x into the table to get the missing equilibrium concentrations

|  | SO$_2$ | O$_2$ | SO$_3$ |
|---|---|---|---|
| Initial | 0.150 | 0.680 | 0.00 |
| Change | -2(0.025) | -0.025 | +2x |
| Eq | 0.100 | 0.655 | 0.05 |

4. Write down the equilibrium expression symbolically and plug in values

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{[0.05]^2}{[0.10]^2[0.655]} = 3.8 \times 10^{-1}$$

- **When the initial concentrations and %dissociation is known:**
-- Ex: What is the Kc for $2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$ if 0.500M of both reactants will be 0.5% dissociated in order to reach equilibrium?

1. Setup the ICE table

|  | SO$_2$ | O$_2$ | SO$_3$ |
|---|---|---|---|
| Initial | 0.500 | 0.500 | 0.0 |
| Change | -2x | -x | +2x |
| Eq | ? | ? | ? |

2. Use the given %dissociation to identify x

$$x = \frac{\% \text{ dissociation}}{100} \cdot [initial\_concentration] \rightarrow x = \frac{0.5}{100} \cdot 0.5 = 0.0025$$

3. Plug x into the table to get the missing equilibrium concentrations

|  | SO$_2$ | O$_2$ | SO$_3$ |
|---|---|---|---|
| Initial | 0.500 | 0.500 | 0.00 |
| Change | -2(0.0025) | -0.0025 | +2x |
| Eq | 0. 4950 | 0.4975 | 0.0050 |

4. Write down the equilibrium expression symbolically and plug in values

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{[0.0050]^2}{[0.4950]^2[0.4975]} = 2.05 \times 10^{-4}$$

- **15.6 Application of Equilibrium Constants**
  - Predicting the Direction of Reaction
    -- when a reaction is not in equilibrium, the mass action expression becomes:
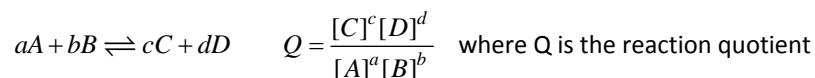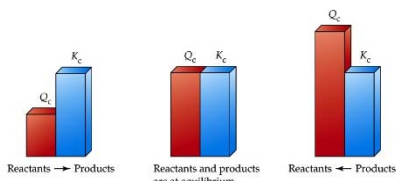
$$aA + bB \rightleftharpoons cC + dD \qquad Q = \frac{[C]^c[D]^d}{[A]^a[B]^b} \quad \text{where Q is the reaction quotient}$$

**Figure: 13.5** from *Chemistry* by McMurray & Fey



-- when Q = K then the reaction is in equilibrium

-- when Q < K then more reactants need to be consumed in order to reach eq

    --- therefore the reaction will proceed in the forward direction

-- when Q < K then more products need to be consumed in order to reach eq

    --- therefore the reaction will proceed in the reverse direction

-- Example: Given the data below is the reaction in equilibrium and if not in
which direction will need to go in order to reach eq?

$A \rightleftharpoons B$    K = 22, [A] = 0.10 M, [B] = 2.0 M
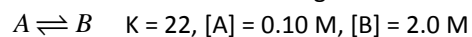
$$Q = \frac{[B]}{[A]} = \frac{2.0}{0.1} = 20 \quad < \quad 22 = K \quad \text{therefore it will go in the forward direction}$$

- Calculating Equilibrium Concentrations

-- Example: The value of $K_c$ = 0.0900 at 298K for the reaction below, determine the eq concentrations if initially [$H_2O$] = 0.00432 M and [$Cl_2O$] = 0.00442 M.

$$H_2O_{(g)} + Cl_2O_{(g)} \rightleftharpoons 2HOCl_{(g)}$$
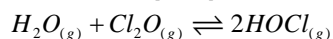
1. *Write down the ICE table*

|        | $H_2O$      | $Cl_2O$     | HOCl  |
|--------|-------------|-------------|-------|
| Initial | 0.00432    | 0.00442     | 0.0   |
| Change | -x          | -x          | +2x   |
| Eq     | 0.00432-x   | 0.00442-x   | +2x   |

2. *Write down the equilibrium expression symbolically*

$$K_c = \frac{[HOCl]^2}{[H_2O][Cl_2O]}$$

3. *Fill in the expression and get the quadratic equation*

$$K_c = \frac{(2x)^2}{(0.00432 - x)(0.00442 - x)} = \frac{4x^2}{1.9094 \times 10^{-5} - 8.74 \times 10^{-2}x + x^2} = 0.0900$$

$$0.0900 \cdot (1.9094 \times 10^{-5} - 8.74 \times 10^{-2}x + x^2) = 4x^2$$

$$1.1785 \times 10^{-6} - 7.866 \times 10^{-4}x + 0.0900x^2 = 4x^2$$

$$0 = 3.9100x^2 + 7.866 \times 10^{-4}x - 1.1785 \times 10^{-6}$$

4. *Use the quadratic equation to find x (choose the positive value)*

Recall the quadratic equation:

$$\text{for} \quad ax^2 + bx + c = 0 \quad x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

in our case, a = 3.9100, b = $7.866 \times 10^{-4}$, c = $1.1785 \times 10^{-6}$

$$x = \frac{-7.866 \times 10^{-4} \pm \sqrt{\left(7.866 \times 10^{-4}\right)^2 - 4 \times 3.9100 \times \left(-1.1785 \times 10^{-6}\right)}}{2 \times 3.9100}$$

$$x = \frac{-7.866 \times 10^{-4} \pm \sqrt{6.187 \times 10^{-7} + 2.688 \times 10^{-5}}}{7.82}$$

$$x = \frac{-7.866 \times 10^{-4} \pm 0.005244}{7.82} \rightarrow x = 5.70 \times 10^{-4} \quad or \quad x = -7.71 \times 10^{-4}$$

5.  *Plug the x value into the E row of the table and find the concentrations:*

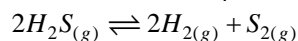x must be greater than zero therefore x = $5.70 \times 10^{-4}$

Using this value we can determine what the concentrations are at eq:

[H$_2$O] = 0.00432 - x = 0.00432 - 0.00057 = 0.00375 M

[Cl$_2$O] = 0.00442 - x = 0.00385 M

[HOCl] = 2x = 0.00114 M

- Example: The value of K$_c$ for the thermal decomposition of hydrogen sulfide is 2.2 x 10$^{-6}$ at 1400K.

$$2H_2S_{(g)} \rightleftharpoons 2H_{2(g)} + S_{2(g)}$$

A sample of gas in which [H$_2$S] = 0.600M is heated to 1400K in a sealed vessel.  After chemical eq has been achieved, what is the value of [H$_2$S]?  Assume no H$_2$ and S$_2$ was present in the original sample.

1.  *Write down the ICE table*

|         | H$_2$S    | H$_2$ | S$_2$ |
|---------|-----------|-------|-------|
| Initial | 0.600     | 0.0   | 0.0   |
| Change  | -2x       | +2x   | +x    |
| Eq      | 0.600-2x  | +2x   | +x    |

2.  *Write down the equilibrium expression symbolically:*

$$K_c = \frac{[H_2][S_2]}{[H_2S]^2}$$

3.  *Fill in the expression*

$$K_c = \frac{(2x)^2(x)}{(0.600 - 2x)^2} = \frac{4x^3}{(0.600 - 2x)^2} = 2.2 \times 10^{-6}$$

4.  *Here we cannot use quadratic so instead we must use an assumption to find x:*

$$assume \ 0.600 \gg 2x \therefore 0.600 - 2x \approx 0.600$$

$$\frac{4x^3}{(0.600 - 2x)^2} = 2.2 \times 10^{-6} \approx \frac{4x^3}{0.36} \rightarrow x = 0.00583M$$

5.  *Verify assumption:*

$$\frac{0.00583 \times 2}{0.600} \times 100\% = 1.94\% < 5\% \therefore valid$$

6.  *Use the E row to find the equilibrium concentrations:*

Using x = 0.00734 M we can determine what the concentrations are at eq:

[H$_2$S] = 0.600 - 2x = 0.600 - 2*0.00583 = 0.588 M

[H$_2$] = 2x = 0.0117 M and [S$_2$] = 0.00583 M

- Example: What are the eq concentrations of each of the species in the following
        reaction, given the $K_c$ = 5.1 at 700K and the initial concentration of all
        species is 0.050 M?          $CO_{(g)} + H_2O_{(g)} \rightleftharpoons CO_{2(g)} + H_{2(g)}$

|          | CO       | H₂O      | CO₂      | H₂       |
|----------|----------|----------|----------|----------|
| Initial  | 0.050    | 0.050    | 0.050    | 0.050    |
| Change   | -x       | -x       | +x       | +x       |
| Eq       | 0.050-x  | 0.050-x  | 0.050+x  | 0.050+x  |

$$K_c = \frac{[CO_2][H_2]}{[CO][H_2O]} = \frac{(0.050+x)(0.050+x)}{(0.050-x)(0.050-x)} = \frac{(0.050+x)^2}{(0.050-x)^2} = 5.1$$

$$\frac{(0.050+x)^2}{(0.050-x)^2} = 5.1 \rightarrow \frac{(0.050+x)}{(0.050-x)} = 2.258$$

$$2.258(0.050-x) = 0.050+x$$

$$0.1129 - 2.258x = 0.050 + x \rightarrow x = 0.0193 \; M$$

using this value we determine the concentrations:
        [CO] = [H₂O] = (0.050-0.0193)M = 0.031 M
        [CO₂] = [H₂] = (0.050+0.0193)M = 0.069 M

- **15.7 Le Châtelier's Principle**
  - defn: when a stressor is applied to a system is at equilibrium, the system will adjust to counteract the
          stressor in order to re-establish eq
  - stressors:
    -- adding or removing reactants or products
    -- changing the pressure/volume
    -- changing the temperature - we will talk about this much later
  - Change in reactant or product concentration, $A \rightleftharpoons B$

    -- if we increase A then $Q = \frac{[B]}{[A]} < K$

        --- we will have to consume A and produce more B to obtain eq
        --- the reaction will go in the forward direction

    -- if we increase B then $Q = \frac{[B]}{[A]} > K$

        --- we will have to consume B and produce more A to obtain eq
        --- the reaction will go in the reverse direction
  - Change in pressure or volume, $2A_{(g)} \rightleftharpoons B_{(g)}$

    -- recall PV = nRT therefore P ∝ n and V ∝ n

    -- if we decrease the volume then $Q = \frac{[B]}{[A]^2} < K$

        --- we increase the number of collisions btwn molecules:
            ---- the pressure increases
            ---- then the no. of moles increases
        --- the system will shift in the direction that reduces the no. of moles
        --- the reaction will go in the forward direction toward product

-- if we increase the volume then $Q = \dfrac{[B]}{[A]^2} > K$

      --- we decrease the number of collisions btwn molecules:
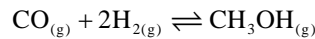
          ---- the pressure decreases

          ---- then the no. of moles decreases

      --- the system will shift in the direction that increases the no. of moles

      --- the reaction will go in the reverse direction toward reactant

- Example: For each scenario predict the direction the reaction goes to attain eq:

$$CO_{(g)} + 2H_{2(g)} \rightleftharpoons CH_3OH_{(g)}$$

    a.) CO is added ……. reaction goes toward product (forward)

    b.) CH₃OH is added ……. reaction goes toward reactants (reverse)

    c.) Pressure is reduced ……. $\Delta n_{reactants}$ = 3, $\Delta n_{products}$ = 1 reaction goes toward reactants (reverse)

    d.) Volume is increased ……. reaction goes toward reactant (forward)

- Change in eq with T

    -- unlike our other changes, the value of the eq const changes with T

    -- if the reaction is exothermic and we add heat then we would predict thereaction would go toward

        reactants: $reactants \rightarrow products + heat$

    -- for an endothermic reaction the addition leads toward more product generation

        $reactants + heat \rightarrow products$

    -- in each case we are treating our heat as a reactant or product

- Example: In what direction will the eq shift when each of the following changes are

        made to the system at eq?

$$N_2O_{4(g)} \rightleftharpoons 2NO_{2(g)} \quad \Delta H^\circ = 58.0 \; kJ$$
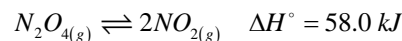
(a) $N_2O_{4(g)}$ is added ……. reaction goes toward product

(b) $NO_{2(g)}$ is removed ……. reaction goes toward product

(c) the total pressure is increased by adding $N_{2(g)}$ ……. reaction remains

        unchanged since the partial pressures of the reacting species is

        constant at constant volume

(d) the volume is decreased ……. reaction goes toward reactant since

      $\Delta n_{reactants}$ = 1, $\Delta n_{products}$ = 2

(e) the temperature is decreased ……. reaction goes toward reactant since

      the process is endothermic or $N_2O_{4(g)} + 58.0 \; kJ \rightleftharpoons 2NO_{2(g)}$

- The Effect of Catalysts – there is no impact on the eq constant it just allows us to reach eq faster