$$\frac{(x)}{(0.100-x)(0.100-x)}=49$$

$$x=49(0.100-x)(0.100-x)$$

$$\downarrow (a-b)^2=a^2-2ab+b^2$$

$$x=49(0.01-0.200x+x^2)$$

$$x=0.49-9.8x+49x^2$$

$$0=49x^2-10.8x+0.49$$

$$a=49 \quad b=-10.8 \quad c=0.49$$

$$x=\frac{10.8\pm\sqrt{(-10.8)^2-4(49)(.49)}}{2(49)}=\frac{10.8\pm\sqrt{20.6}}{98}$$

$$x=\cancel{0.157} \text{ or } 0.0639$$

$$x=0.0639$$

The QUADRATIC EQUATION:

$$ax^2+bx+c=0$$

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

Each quadratic has two solutions (see the +/- part of the equation), but onlu one of them will be the correct chemical solution.

We will discard the value of x=0.157, since this would result in NEGATIVE concentrations of phosphorus trichloride and chlorine gas ... which is physically impossible.

| Species | [Initial] | $\Delta$ | [Equilibrium] |
|---|---|---|---|
| $PCl_3$ | $\frac{0.400 \text{ mol}}{4.00 \text{ L}} = 0.100 M$ | $-X$ | $0.100 - X$ |
| $Cl_2$ | $\frac{0.400 \text{ mol}}{4.00 \text{ L}} = 0.100 M$ | $-X$ | $0.100 - X$ |
| $PCl_5$ | 0 | $+X$ | $X$ |

$X = 0.0639$

EQUILIBRIUM CONCENTRATIONS

Moles of each species at equilibrium

$[PCl_3] = 0.100 - x = 0.036 M \times 4.00 L = 0.144 \text{ mol } PCl_3$

$[Cl_2] = 0.100 - x = 0.036 M \times 4.00 L = 0.144 \text{ mol } Cl_2$

$[PCl_5] = x = 0.0639 M \times 4.00 L = 0.256 \text{ mol } PCl_5$

Quick comparison of initial and equilibrium states:

$$\boxed{\begin{array}{l} 0.400 \text{ mol } PCl_3 \\ 0.400 \text{ mol } Cl_2 \\ 0 \text{ mol } PCl_5 \end{array}} \xrightarrow{K_c = 49} \boxed{\begin{array}{l} 0.144 \text{ mol } PCl_3 \\ 0.144 \text{ mol } Cl_2 \\ 0.256 \text{ mol } PCl_5 \end{array}}$$

114 An 8.00 L reaction vessel at 3900C is charged with 0.850 mol of nitrogen and oxygen gases. Find the concentration of all species at equilibrium.

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g) \quad K_c = 0.0123$$

$$K_c = \boxed{0.0123 = \frac{[NO]^2}{[N_2][O_2]}}$$

To solve this, we must express all these equilibrium concentrations in terms of one variable.

| Species | [Initial] | Δ | [Equilibrium] |
|---|---|---|---|
| $N_2$ | $\frac{0.850\,mol}{8.00\,L} = 0.10625\,M$ | $-x$ | $0.10625 - x$ |
| $O_2$ | $\frac{0.850\,mol}{8.00\,L} = 0.10625\,M$ | $-x$ | $0.10625 - x$ |
| NO | 0 | $+2x$ | $2x$ |

Defined 'x' as the change in nitrogen concentration.

$$\frac{[NO]^2}{[N_2][O_2]} = \boxed{\frac{(2x)^2}{(0.10625-x)(0.10625-x)} = 0.0123}$$

We must solve this expression for 'x' to solve the problem.

$$\frac{(2x)^2}{(0.10625-x)(0.10625-x)} = 0.0123$$

$$\sqrt{\frac{(2x)^2}{(0.10625-x)^2}} = \sqrt{0.0123}$$

$$\frac{2x}{0.10625-x} = 0.110905365$$

$$2x = 0.110905365(0.10625-x)$$

$$18.0333926x = 0.10625-x$$

$$19.0333926x = 0.10625$$

$$x = 0.005582943$$

| Species | [Equilibrium] |
|---|---|
| $N_2$ | $0.10625-x$ |
| $O_2$ | $0.10625-x$ |
| NO | $2x$ |

After finding 'x', plug back into the table to find each concentration!

$N_2$: $0.10625-x = \boxed{0.101\ M\ N_2}$

$O_2$: $0.10625-x = \boxed{0.101\ M\ O_2}$

NO: $2x = \boxed{0.0112\ M\ NO}$

Since Kc is small (0.0123), we expect that reactants will dominate at equilibrium.

That's what we see
<--- here!

# PRESSURE AND EQUILIBRIUM

- Pressure can affect a GAS-PHASE equilibrium ... sometimes. How?

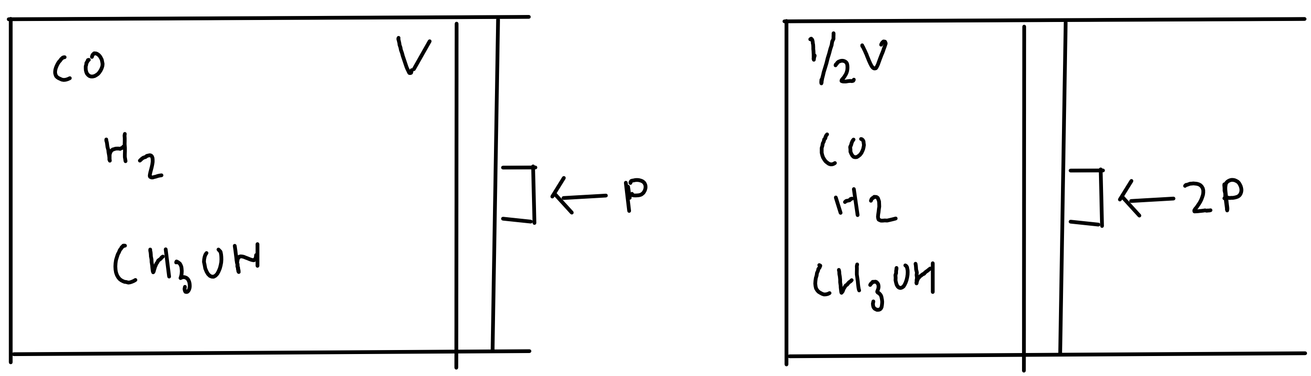
$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

... how might pressure affect this equilibrium?

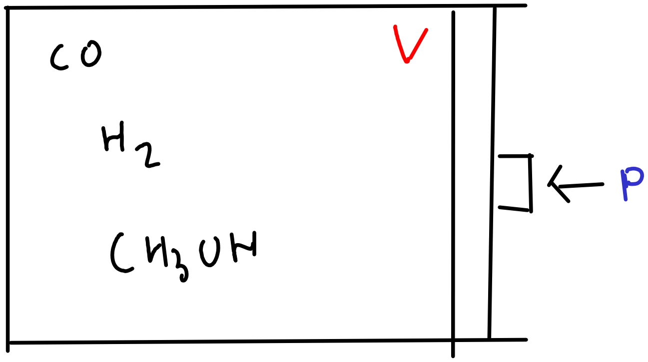
- If the change in pressure CHANGES CONCENTRATIONS, then this equilibrium would be disturbed and Le Chateleir's Principle would apply.

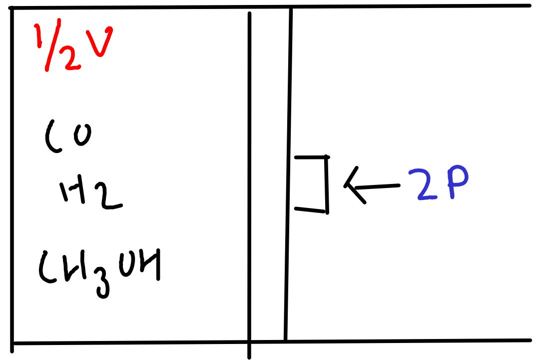
    - Adding an INERT GAS would change pressure, but would it change concentration of the gases? NO - so addition of argon would have no effect on the equilibrium!

        - What about COMPRESSION?




... compression increases pressure by DECREASING total volume.




... but this volume change affects ALL concentrations the same way. In this example, each concentration is DOUBLED.

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

$$K_c = \frac{[CH_3OH]}{[CO][H_2]^2} \qquad \frac{(1)}{(1)(1)^2} = 1$$

For simplicity, let's assume Kc = 1, and all concs = 1M

Doubling concentrations gives Q= $\dfrac{2}{(2)(2)^2} = \dfrac{1}{4}$

$Q < K_c$, so equilibrium shifts to the RIGHT, forming more methanol at the expense of hydrogen and carbon monoxide.

In general, compressing an equilibrium reaction in the gas phase will cause the equilibrium to shift towards the side with fewer moles of gas. This causes the pressure to decrease.

In general, decompressing an equilibrium reaction in the gas phase will cause the equilibrium to shift towards the side with more moles of gas. This causes the pressure to increase.

HOWEVER, this can only be true IF there's a side of the reaction with more moles of gas than the other. If both sides of the reaction have the SAME number of moles of gas, then a pressure change will NOT affect the equilibrium.

Example: $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$

... would not respond to a pressure change.

# FACTORS THAT MAY AFFECT EQUILBRIUM

(1) TEMPERATURE (effect depends on whether reaction is endothermic or exothermic)

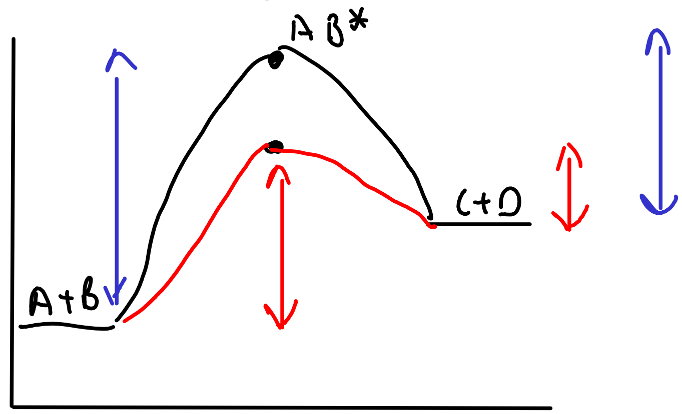
- Changes rate of reaction, too!

... changes Kc

(2) PRESSURE - only for gas-phase reactions which have different numbers of moles of gas on each side of the equilbrium.  Otherwise, no effect.

... no change of Kc

(3) CATALYSTS - do NOT affect equilibrium, but make the equilbrium state occur more quickly.




The catalyst raises BOTH forward and reverse rates, so it doesn't affect the composition of the equilibrium mixture!

(4) CONCENTRATION - Le Chateleir's Principle applies for changing concentrations.  An equilibrium will shift to counteract a change in concentration of reactant or product.

... doesn't change Kc.

# ACID/BASE EQUILIBRIUM

- Several scientific theories exist that define acid-base chemistry. We will discuss THREE of these theories.

- These theories differ in the way that acids, bases, and their associated reactions are defined.

- Typically, the newer theories include MORE chemicals under the umbrella of "acid-base chemistry"!

## THREE ACID-BASE THEORIES

1. Arrhenius theory

2. Bronsted-Lowry theory

3. Lewis theory