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.

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

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

We need to express all the concentrations in terms of a single variable.

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

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

We need to solve this expression for 'x' to find our conecntrations.

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

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

You may solve this problem using either the quadratic formula or, (easier) taking the square root of both sides.

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

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

$$2x = 0.011783695 - 0.110905365x$$

$$2.110905365x = 0.011783695$$

$$x = 0.0055822943$$

Now, we use 'x' to find the equilibrium concentrations:

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

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

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

Remember, Kc = 0.0123

We expect reactants to dominate at equilibrium

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

PRESSURE AND EQUILIBRIUM

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

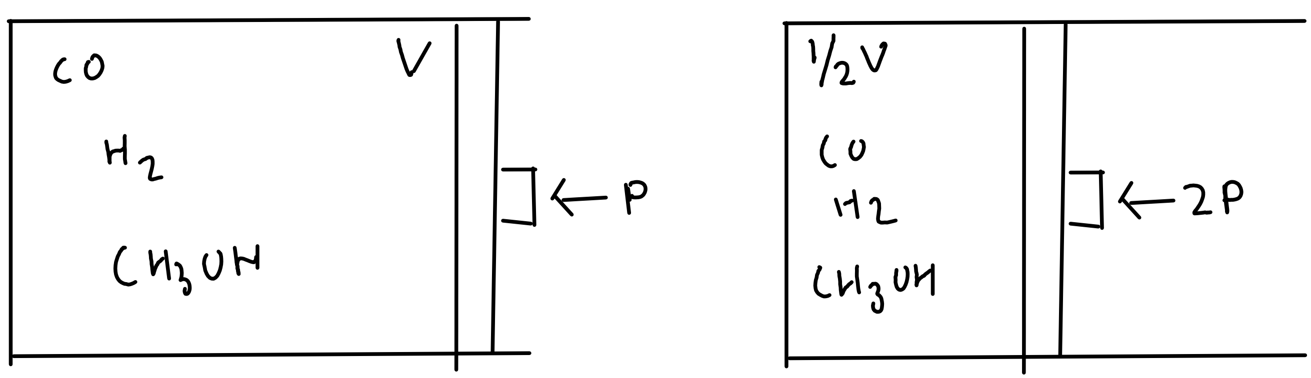
$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

... how might pressure affect this equilibrium?

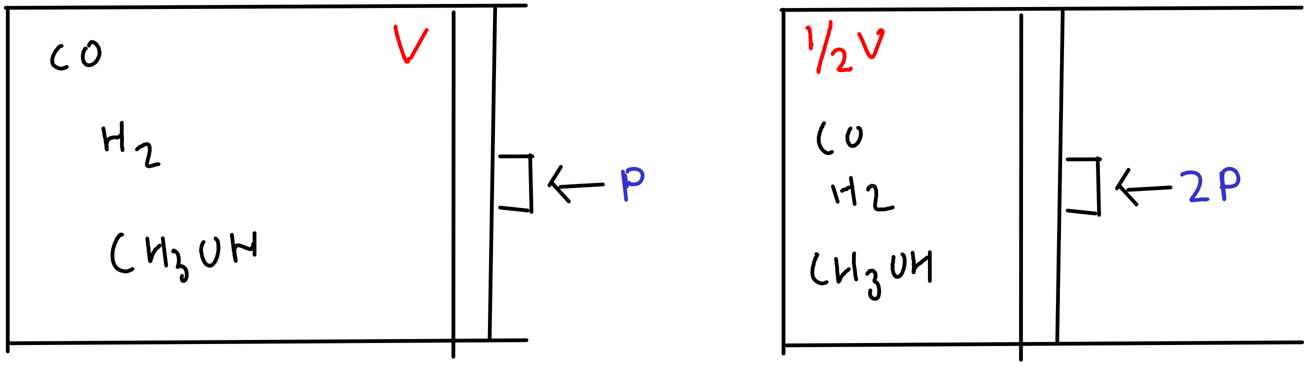
- If the change in pressure CHANGES CONCENTRATIONS, then this equilibrium would be disturbed and Le Chateleir's Principle would apply.

    - 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= $\frac{2}{(2)(2)^2} = \frac{1}{4}$

$Q < K_c$, 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)

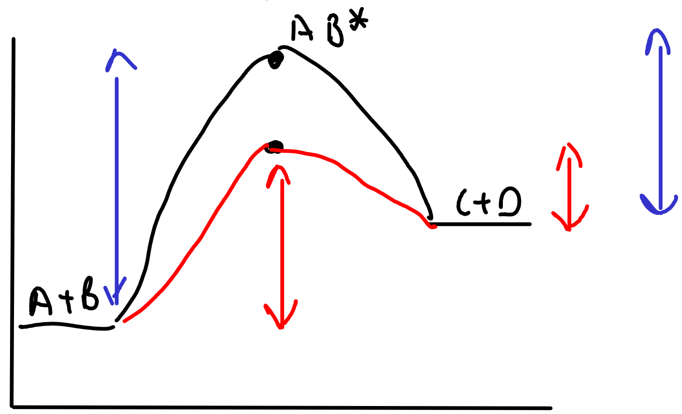
   - Changes rate of reaction, too!

   ... 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.