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.

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

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

We need to express these EQUILIBRIUM concentrations in terms of a single variable.

SPECIES	INITIAL CONCENTRATION	CHANGE	EQUILIBRIUM CONNCENTRATION
N_2	$\frac{0.850 \text{ mol}}{8.00 L} = 0.10625 M$	$-x$	$0.10625 - x$
O_2	$\frac{0.850 \text{ mol}}{8.00 L} = 0.10625 M$	$-x$	$0.10625 - x$
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 equilibrium concentrations.

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

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

Solve by taking square root of both sides

SPECIES	EQUILIBRIUM CONNCENTRATION
N_2	$0.10625 - x$
O_2	$0.10625 - x$
NO	$2x$

$$\frac{2x}{(0.10625-x)} = 0.1109053651$$

$$2x = 0.011783695 - 0.1109053651x$$

$$2.1109053651x = 0.011783695$$

$$x = 0.005582$$

Now, calculate the concentrations by plugging 'x' into our expressions

$$[N_2] = .10625 - x = 0.101\ M$$

$$[O_2] = .10625 - x = 0.101\ M$$

$$[NO] = 2x = 0.0112\ M$$

Since Kc = 0.0123, we expect that the reactants will dominate at equilibrium

PRESSURE AND EQUILIBRIUM

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

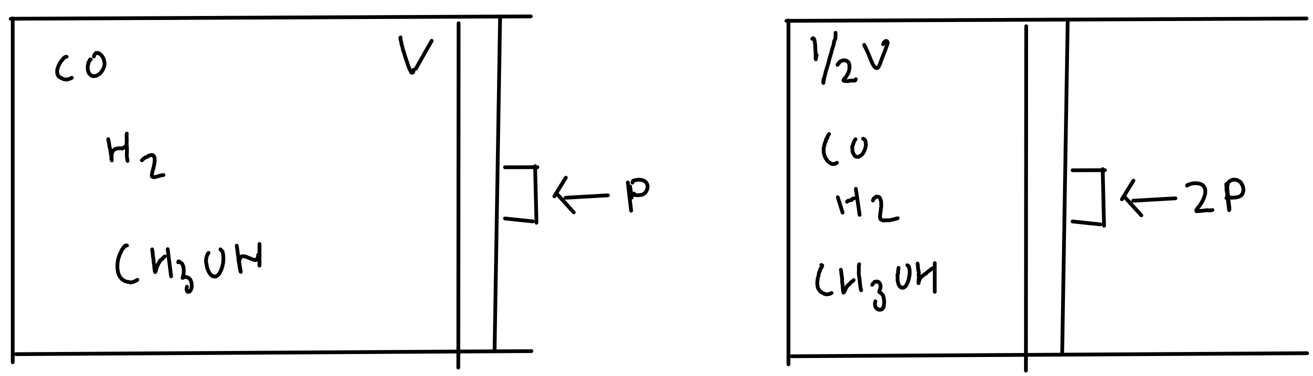
$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

... how might pressure affect this equilibrium?

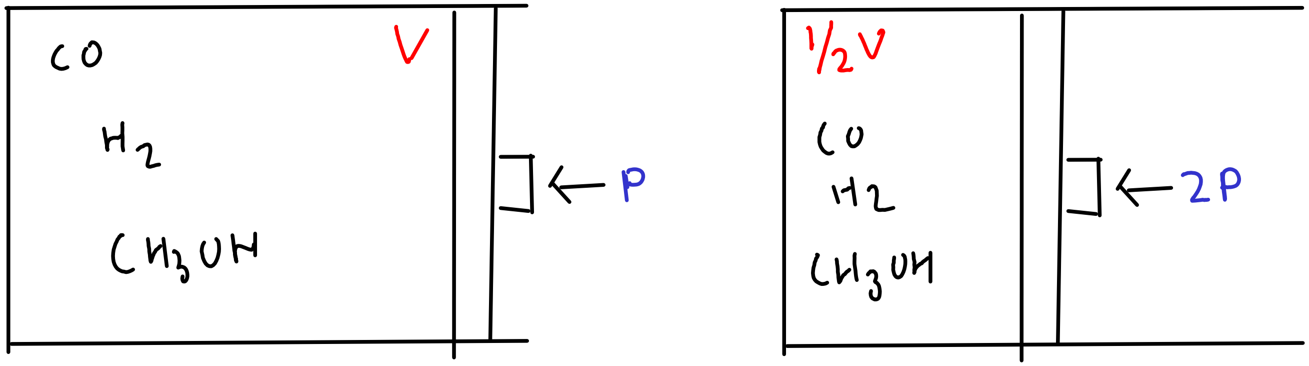
- If the change in pressure CHANGES CONCENTRATIONS, then this equilibrium would be disturbed and Le Chateleir's Principle would apply.

  - 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)

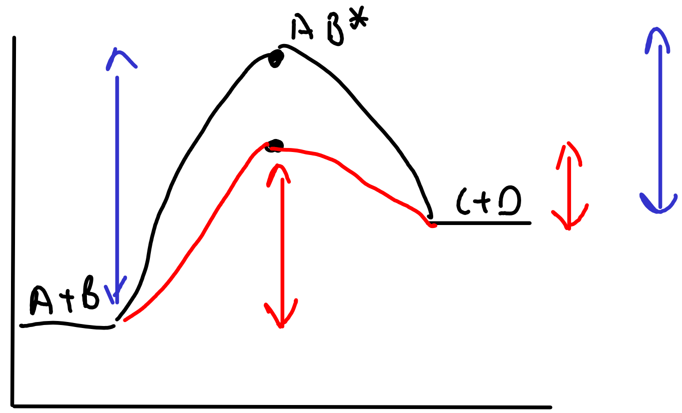
- Changes rate of reaction, too!

... 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 be reached faster.



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

ARRHENIUS THEORY

- The oldest model of acid-base chemistry!

- Only applicable to systems where WATER is the solvent!

ACIDS are substances that ionize in water to increase the concentration of HYDRONIUM ION

$$HA + H_2O \rightleftharpoons \underbrace{H_3O^+}_{\text{Hydronium ion}} + A^-$$

or, for simplicity: $HA \overset{H_2O}{\rightleftharpoons} H^+ + A^-$

"Hydrogen ion" - doesn't really exist as a free ion in water, but a convenient simplification!

ARRHENIUS THEORY

BASES are substances that ionize in water to increase the concentration of HYDROXIDE ION

For soluble metal hydroxides:

$$NaOH \rightarrow Na^+ + OH^-$$

$$MOH \xrightleftharpoons{H_2O} M^+ + \underbrace{OH^-}_{\text{Hydroxide ion}}$$

For other Arrhenius bases:

$$B + H_2O \rightleftharpoons BH^+ + OH^- \quad \text{ex: } NH_3$$

An Arrhenius acid base reaction can be represented by:

$$\boxed{H_3O^+ + OH^- \rightleftharpoons 2H_2O}$$

"neutralization"

or, using hydrogen ion instead of hydronium

$$H^+ + OH^- \rightleftharpoons H_2O$$

BRONSTED-LOWRY THEORY

H^+ ions!

- Bronsted-Lowry theory views acid-base reactions as PROTON TRANSFER reactions!

ACIDS are PROTON DONORS

BASES are PROTON ACCEPTORS

$$HCl + NH_3 \rightleftharpoons NH_4^+ + Cl^-$$

acid, base, conjugate acid of NH_3, conjugate base of HCl

A CONJUGATE PAIR is an acid and a base that differ by a proton!