1. Chemical equilibrium: various definitions:
   - $\Delta G_{\text{rxn}} = 0 \rightarrow$ most general formula = $\Delta G_{\text{products}} - \Delta G_{\text{reactants}}$
     → using $\Delta G^o = \Delta H^o - T\Delta S^o + RT \ln p$; where $^o$ means standard conditions (25 °C and 1 atm pressure), $T\Delta S^o$ is the temperature- and $RT \ln p$ is pressure-dependence
     - \[ \text{[Co(OH}_2)_6]^{2+} + 4 \text{Cl}^- \rightleftharpoons [\text{CoCl}_4]^{2-} + 6 \text{H}_2\text{O} \]
     - $\Delta H^o > 0$ kJ/mol, $\Delta S^o > 0$ kJ/mol, $\Delta G(T) < 0$
     - at $^o$ conditions violet solution, cool down becomes pink, warm up blue
     → changing $T,p$ will shift the reaction toward a new equilibrium

   - A state where the concentration of products and reactants remain constant.
     → consider adding AgNO$_3$ or NaCl to the above reaction
     → remember dissolving CuCl$_2$ in water than adding ammonia than adding HCl gas ...
     → adding or removing products or reactants will shifts the reaction to a new equilibrium

   - A dynamic situation where the forward rate/speed of the reaction equals with the reverse rate
     - $\text{H}_2\text{O}(g) + \text{CO}(g) \rightleftharpoons \text{H}_2(g) + \text{CO}_2(g)$
     → introduce a change (↓ temperature = slow down the molecules, ↑ pressure = increase the speed of the molecules ~ collisions) to a closed container and reaction moves to a new equilibrium

2. Characterization of chemical equilibrium with equilibrium constant ($K$)
   - consider ammonia synthesis:
     - $\text{N}_2(g) + 3 \text{H}_2(g) \rightleftharpoons 2 \text{NH}_3(g)$

   - $t = 25 ^\circ \text{C}$ ($^o$ normal conditions) $\Delta H^o = -92 \text{ kJ/mol}, \Delta S^o = -34 \text{ J/molK}, \Delta G^o = -34 \text{ kJ/mol}$;
     - since $\Delta G^o < 0$ this is a spontaneous process and not at the equilibrium $\Rightarrow$ large kinetic barrier (need high collision energy to activate the inert $\text{N}_2$ and $\text{H}_2$ molecules) $\rightarrow$ reaction is kinetically controlled

   - the law of mass action is used to describe the equilibrium conditions:
     - in general for a reaction $j \text{A} + k \text{B} \leftrightarrow l \text{C} + m \text{D}$ then the equilibrium constant is:
       \[ \frac{[C]^l[D]^m}{[A]^j[B]^k} \]
     - by convention $K$ is dimension less; however, it has $(\text{mol/dm}^3)^{ΔN}$

   - $t = 127 ^\circ \text{C}$; there is some limited formation of the ammonia:
     \[ [\text{NH}_3] = 3.1 \cdot 10^{-2} \text{ mol/dm}^3; [\text{N}_2] = 8.5 \cdot 10^{-1} \text{ mol/dm}^3; [\text{H}_2] = 3.1 \cdot 10^{-3} \text{ mol/dm}^3 \Rightarrow K = 3.8 \cdot 10^4 \]
   - for the reverse reaction $2\text{NH}_3 \leftrightarrow \text{N}_2 + 3 \text{H}_2$ $K_{\text{reverse}} = K^{-1}$

   - $t = 500 ^\circ \text{C}$; $K = 6 \cdot 10^2$ $\rightarrow$ strong temperature dependence of $K$ $\rightarrow$ can we analytically determine this?

3. Temperature dependence of the equilibrium constant
   - Remember from free-energy expression in equilibrium $\Delta G^o = -RT \ln K \Rightarrow \Delta H^o - T\Delta S^o = -RT \ln K$
   - $\ln K(T) = - \frac{\Delta G^o}{RT} = - \frac{\Delta H^o}{RT} (1/T) + \frac{\Delta S^o}{R}$
   - Does this expression look familiar? What was the constant before?
4. Pressure dependence of the equilibrium constant (remember: concentration in gas phase = partial pressure) (remember: vapor pressure of mixed liquids = \( \Sigma \chi_i p_{\text{vap},i} \))

- consider the ideal gas equation \( p = \frac{n}{V} RT = c RT \) (remember: osmotic pressure)
  where \( c \) is the molar concentration of the gas (molarity)

- \( K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} \) or \( K_p = \frac{p_{\text{NH}_3}^2}{p_{\text{N}_2} p_{\text{H}_2}^3} \) → interconversion by using \( c = p/RT \) gives \( K_p = K_c (RT)^{\Delta N} \),
  where \( \Delta N \) is the difference between the number of particles of the product and reactant sides