

HL Topic 7 Equilibrium - Guided Notes

Slide 3: ICE Tables

- Used to calculate equilibrium constants or equilibrium concentrations when the equilibrium concentrations of all reactants and/or products are not given.
- Remember to make sure the precision of any answer to calculations is consistent with the data given in the question (aka use CORRECT SIG FIGS)
- ICE
- initial
- change
- equilibrium

Slide 4: Guidelines for ICE Tables

- Write _____ chemical equation
- Under the equation, write three rows labeled *initial*, *change*, and _____.
- *Initial*: _____ originally placed in the flask. Unless otherwise indicated, assume initial product concentration is zero.
- *Change*: represents the amount that reacts to reach equilibrium. A minus sign represents a decrease in _____ of reactants. A plus sign represents an increase in concentration of products.
- *Equilibrium*: _____ in the equilibrium mixture.
- Equilibrium _____ = initial concentration \pm change in concentration
- Write the equilibrium expression and calculate K_c from the equilibrium _____.

Slide 5: $\text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons \text{PCl}_5(\text{g})$

- ICE Table: Calculating K_c from initial and equilibrium _____
- _____ amount PCl_3 : 0.20 mol
- _____ amount Cl_2 : 0.10 mol
- _____ amount PCl_3 : 0.12 mol
- Volume flask: 1.0 dm³
- Initial _____ PCl_3 : 0.20 mol dm⁻³
- Initial _____ Cl_2 : 0.10 mol dm⁻³
- Equilibrium _____ PCl_3 : 0.12 mol dm⁻³
- $\text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons \text{PCl}_5(\text{g})$
- ***Remember to divide amount by volume to get _____**
- 0.20

- 0.10
- 0.00
- **-0.08**
- **-0.08**
- **+0.08**
- 0.12
- **0.02**
- **0.08**
- Initial (mol dm⁻³)
- Change (mol dm⁻³)
- _____ (mol dm⁻³)
- $K_c = \frac{[\text{PCl}_5]}{[\text{PCl}_3][\text{Cl}_2]}$
- = 0.08
- (0.12)(0.02)
- = 33
- This change is also 0.08 because PCl₃ and Cl₂ react in a 1:1 _____ ratio

Slide 6: Let's Practice!

- The oxidation of NO to form NO₂ occurs during the formation of smog. When 0.60 mol of NO was reacted with 0.60 mol of O₂ in a 2.0 dm³ container at 500°C, the equilibrium mixture was found to contain 0.20 mol of NO₂. Calculate the equilibrium constant for the reaction at this temperature.

Slide 7: CO(g) + 2H₂(g) ⇌ CH₃OH(g)

- Equilibrium _____ CO: 0.200 mol dm⁻³
- Equilibrium _____ H₂: 0.155 mol dm⁻³
- $K_c = \frac{[\text{CH}_3\text{OH}]}{[\text{CO}][\text{H}_2]^2}$
- Calculating equilibrium _____ from the equilibrium constant
- $K_c = 0.500$ at 350K.
- _____
- $[\text{CH}_3\text{OH}] = \frac{0.500 \times (0.200)(0.155)^2}{0.500} = 0.00240 \text{ mol dm}^{-3}$ or $2.40 \times 10^{-3} \text{ mol dm}^{-3}$

Slide 8: SO₃(g) + NO(g) ⇌ NO₂(g) + SO₂(g)

- Change in concentration NO = -x

- Change in concentration $\text{SO}_3 = -x$
- Change in concentration of NO_2 and $\text{SO}_2 = +x$
- $6.78 = [\text{NO}_2][\text{SO}_2]$
- $(0.0300-x)^2$
- ICE Table: Calculating K_c from initial concentrations
- The equilibrium constant for the reaction was found to be 6.78 at a specified temperature. If the initial concentrations of NO and SO_3 were both $0.0300 \text{ mol dm}^{-3}$, what would be the equilibrium concentration of each component?
- $\text{SO}_3(\text{g}) + \text{NO}(\text{g}) \rightleftharpoons \text{NO}_2(\text{g}) + \text{SO}_2(\text{g})$
- ***Use algebra to calculate equilibrium concentrations.**
- 0.0300
- 0.00
- **-x**
- Initial (mol dm^{-3})
- Change (mol dm^{-3})
- Equilibrium (mol dm^{-3})
- $[\text{SO}_3][\text{NO}]$
- $= x^2$
- **1:1 stoichiometric ratios between reactants and products**
- **0.0300-x**
- 0.0300
- **-x**
- **0.0300-x**
- 0.00
- **+x**
- **+x**
- **x**
- **x**
- $\sqrt{6.78}$
- $= x$
- $0.0300-x$
- $2.60(0.0300-x) = x$
- $0.0780 - 2.60x = x$
- $x = 0.0217$
- $[\text{SO}_3] = [\text{NO}] = 0.0300 - 0.0217 = 0.00830 \text{ mol dm}^{-3}$
- $[\text{NO}_2] = [\text{SO}_2] = 0.0217 \text{ mol dm}^{-3}$

Slide 9: What about when K_c is very small?

- When K_c is small ($< 10^{-3}$), this means very little product is made

- The equilibrium mixture contains mostly reactants.
- The change in reactant concentrations is close to zero.
- Assume initial reactant concentrations are approximately the same as equilibrium concentrations.

Slide 10: $2\text{H}_2\text{O}(\text{g}) \rightleftharpoons 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$

- $(0.10)^2$
- $(2x)^2(x)$
- $= (2x)^2(x)$
- Change in _____ $\text{H}_2 = 2x$
- Change in _____ $\text{O}_2 = x$
- Change in _____ of $\text{H}_2\text{O} = -2x$
- $K_c = [\text{H}_2]^2[\text{O}_2]$
- $(0.10)^2$
- ICE Table: Calculating equilibrium _____ when K_c is very small
- The thermal _____ of water has a very small value of $K_c = 7.3 \times 10^{-18}$ at 1000°C . A reaction is set up at this temperature with an initial H_2O concentration of 0.10 mol dm^{-3} . Calculate the H_2 concentration at equilibrium.
- $2\text{H}_2\text{O}(\text{g}) \rightleftharpoons 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$
- 0.10
- 0.00
- **-2x**
- Initial (mol dm^{-3})
- Change (mol dm^{-3})
- _____ (mol dm^{-3})
- $[\text{H}_2\text{O}]^2$
- **2:1** _____ ratio between H_2O and O_2 and between H_2 and O_2
- **0.10-2x**
- **2x**
- 0.00
- **+2x**
- **+x**
- **x**
- $7.3 \times 10^{-18} =$
- $4x^3 = (7.3 \times 10^{-18})(0.10)^2$
- $4x^3 = 7.3 \times 10^{-20}$
- $[\text{H}_2] = 2x = 2(2.632 \times 10^{-7}) = 5.3 \times 10^{-7} \text{ mol dm}^{-3}$
- $x = 2.632 \times 10^{-7}$
- $[\text{H}_2\text{O}]_{\text{initial}} \approx [\text{H}_2\text{O}]_{\text{_____}}$

- = 0.10

Slide 11: Practice Problems

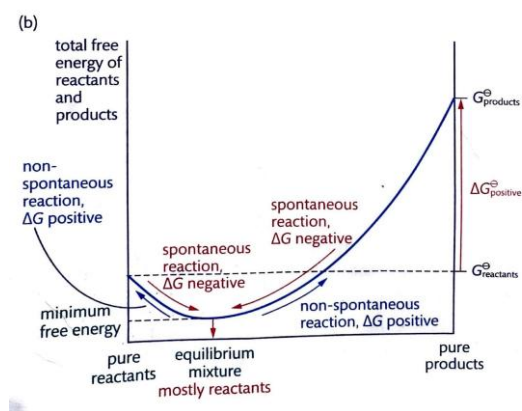
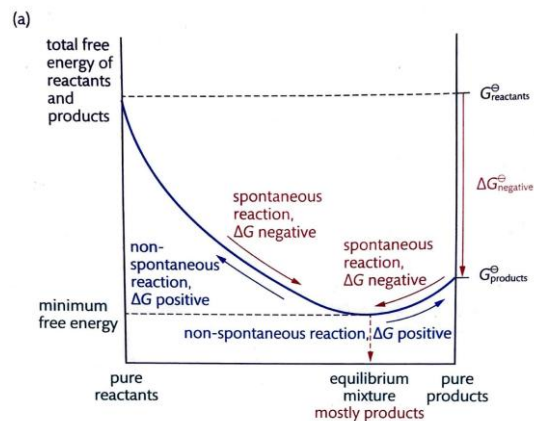
- The dissociation of hydrogen iodide into its elements takes place in a 1.0 dm³ container at 440°C. When 1.0 mol of hydrogen iodide is used, it is found to have decreased to 0.78 mols at equilibrium.
- Calculate the equilibrium constant for this reaction at this temperature.
- Deduce whether the dissociation reaction is endothermic or exothermic, given at 600°C the value of K_c is 0.04.
- The reaction $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$ is carried out in a closed container with initial concentrations of both reactants of 1.6 mol dm⁻³. K_c for the reaction is 1.7×10^{-3} . Calculate the concentration of NO(g) at equilibrium.
- (a) The reaction $\text{CO}(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{CO}_2(\text{g})$ was studied at 550°C. When 4.0 mol each of CO and H₂O were introduced into a 1 dm³ vessel, the equilibrium mixture was found to contain 3.2 mol of both H₂ and CO₂. Calculate the concentrations of CO and H₂O at equilibrium and the value of K_c .
- (b) At the same temperature and pressure, a different experiment was found to have 4.0 mol of both CO and H₂O, and 3.0 mol of both H₂ and CO₂ present in the mixture after a period of reaction time. Show mathematically that this mixture had not reached equilibrium, and in which direction it will react.

Slide 12: Free Energy and Equilibrium

- Why do some reactions have more products than reactants at equilibrium and vice versa?
- Has to do with the Gibbs free energy, ΔG (aka work available), from a system.
- The more Gibbs free energy available, the more the reaction will proceed towards the products.
- As a reaction progresses, the free energy of the reactants decreases and the free energy of the products increases.
- At equilibrium, the total free energies of the reactants and products are equal.
- $\Delta G_{\text{products}} = \Delta G_{\text{reactants}}$
- Equilibrium occurs at the **lowest energy state**
- More negative ΔG means more work available for the reaction

Slide 13: Gibbs free energy

- $\Delta G = \Delta G_{\text{products}} - G_{\text{reactants}}$
- ΔG is negative for a reaction that proceeds in the forward direction (a reaction that is _____)
- $\Delta G < 0$
- $\Delta G > 0$



Slide 14: Summary

- ΔG

Slide 15: Calculation of K_c from Thermodynamic Data

- The _____ constant is related to Gibbs free energy through the following equation:
- $\Delta G = -RT \ln K$
- _____ free energy change of rxn
- Gas _____ ($8.31 \text{ J K}^{-1} \text{ mol}^{-1}$)
- Temperature (K)
- Natural log of _____ constant (K_c)

Slide 16: Relationship between ΔG and K_c

- $\Delta G = -RT \ln K$

Slide 17: Calculations with free energy

- This equation can be used when the equilibrium constant can't be measured correctly
- When the reaction is too slow to reach equilibrium

- When amounts of reactants/products are too small to measure directly.
- $\Delta G = -RT \ln K$
- **Example:**
- The esterification reaction that produces ethyl ethanoate has a free energy change $\Delta G = -4.38 \text{ kJ mol}^{-1}$
- $\text{CH}_3\text{COOH (aq)} + \text{C}_2\text{H}_5\text{OH (aq)} \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5\text{(aq)} + \text{H}_2\text{O(aq)}$
- Calculate the value of the equilibrium constant of this reaction at 298K.

Slide 18: Kinetics

- Equilibrium



Slide 19: Kinetics and Equilibrium

- The magnitude of the equilibrium constant gives no information about the rate of the reaction.
- Example: the rusting of iron has a $\Delta G = -1490 \times 10^6 \text{ J}$ and $K_c = 10^{261}$
- The large K_c means the reaction is thermodynamically favorable and will proceed towards completion. A large, negative ΔG means the reaction is spontaneous. However, rusting is a slow process that can take years.

Slide 20: At equilibrium: $k[A][B] = k'[C][D]$

Rearranging this equation:

k

- Rate constants of forward and reverse reactions can help explain some aspects of _____.
- $A + B \rightleftharpoons C + D$
- Given:
- Rate of forward _____ = $k[A][B]$
- Rate of reverse _____ = $k'[C][D]$
- $K_c = \frac{[C][D]}{[A][B]}$
- _____ between rate and equilibria
- k'
- $\frac{[A][B]}{[C][D]}$
- $\frac{[A][B]}{[C][D]}$
- $K_c = \frac{k}{k'}$
- k'
- k
- $k \gg k'$ K_c is large and _____

Slide 21: Let's Practice!

Slide 22: Turn and Talk

- The study of chemical change often involves both the _____ and microscopic scales. Compare what is observed at both the macroscopic and microscopic levels of a reaction once equilibrium is reached.