

Using the “ICE” Method to Calculate K, Given Initial Concentrations and One Final Concentration

1. Write balanced equation, and expression for K_c
2. Make an “ICE” table, and enter the knowns
 - a. Initial
 - b. Change
 - c. Equilibrium
3. Find the Change for the chemical whose final concentration is known
4. Use stoichiometric relationship to determine the change in concentrations for the others
5. From the initial concentrations and the deduced changes, determine all equilibrium concentrations
6. With all equilibrium concentrations now known, plug into the K_c expression and solve for K
7. Check: Does Answer Make Any Sense?
 - Note: equilibrium concentrations must be in Molarity, moles/liter. If information is given in grams or moles plus solvent volume, you will need to convert into molarity.

Using the “ICE” Method to Calculate Equilibrium Concentrations, Given Only Initial Concentrations and K_c

1. Write balanced equation, and expression for K_c
2. Make an “ICE” table, and enter the known initial concentrations
 - a. Initial
 - b. Change
 - c. Equilibrium
3. Use “x” to define the change of one substance.
4. Use stoichiometric relationships to determine the changes in the concentrations for the others, in terms of “x”.
5. Calculate the equilibrium concentrations of all chemicals in terms of initial concentrations and “x”, and enter them in the table.
 - Ex: $0.30 - x$, or $0.30 - 2x$, or $0.00 + x$, or $0.00 + 2x$
 - **If K is small** so that “x” is likely to be small, **use the simplifying assumption** that $[A]_{\text{initial}} - “x” = [A]_{\text{initial}}$
 - This is often justified, and can greatly simplify the math.
 - Ex: $0.20 - x = 0.20$ if x is smaller than 0.01
6. Solve for “x” (This is the hard part!)
7. **Once “x” is known, use it to solve for the actual equilibrium concentrations**
8. Check: Does Answer Make Any Sense?
9. Check: If you made the “simplifying assumption”, was it justified? (Was “x” $< 5\%$ of $[A]_{\text{initial}}$?)