

# CHE 1302

## Basic Principles of Modern Chemistry II

### Week 6

Hello and welcome to the weekly resources for Chemistry 1302! This resource covers topics typically taught by professors during the 6<sup>th</sup> week of classes.

On our website, <https://baylor.edu/tutoring>, you'll find the following links:

"Online Study Guide Resources" – If you don't see the topics you're learning right now, click here to find the weekly resources for the rest of the semester!

"How to Participate in Group Tutoring" - See if there is a Chemistry 1302 group tutoring session being hosted this semester – these are weekly question/answer sessions taught by our master tutors!

You can also view tutoring times for your course or schedule a private 30-minute appointment! Check out the website to learn more. You can also give us a call at (254)710-4135, or drop in. Our hours are Monday-Thursday 9 am – 8 pm on class days.

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**KEY WORDS: ICE Tables, Equilibrium Calculations, X is Small**

## TOPIC OF THE WEEK:

### ICE Tables

How to solve problems involving a disturbed equilibrium state:

The following device is called a RICE table (you may also hear it called an ICE table). It is a helpful way to keep track of the changes for each reactant/product!

Reaction	Write the <b>equation</b> here. You will not do any of the calculations below on solids/liquids!
Initial concentrations	Beneath each reactant and product, write its <b>concentration at the start</b> of the experiment. If the problem doesn't say that any was added/present, assume initial concentration = 0.
Change *here, "change" refers to the change in concentration between the	<ol style="list-style-type: none"><li>1. Write <b>ONE change in concentration</b>.<ol style="list-style-type: none"><li>a. If you know that some reagent gets used up completely: the change in concentration in that reagent will be equal to its initial concentration.</li><li>b. <b>If you don't know how much the reagents changed, write "x"</b> under any one of them.</li></ol></li><li>2. Fill in the rest of the changes in concentrations based on that one.</li></ol>

start of a reaction and when it reaches equilibrium	<p>a. These changes are related stoichiometrically. Multiply each reagent's change by this ratio: <b>(its coefficient)/(coefficient of the reagent from step 1)</b></p> <p>b. To figure out the signs of each "change," calculate Q and compare it to K.</p> <p>i. If <math>Q &lt; K</math>, the ratio of products/reactants will want to increase to get to equilibrium. So, the "changes" on the product side will be positive, and the "changes" on the reactant side will be negative.</p> <p>ii. If <math>Q &gt; K</math>, the ratio of products/reactants will want to decrease to get to equilibrium. So, the "changes" on the product side will be negative, and the "changes" on the reactant side will be positive.</p>
Equilibrium	Add the I and C rows together.

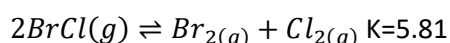
## Highlight 1: Problem-Solving: ICE Tables

Different problems will start you with different pieces of information, but you can consistently follow these steps:

1. Write the **balanced chemical equation**.
2. Write **K**'s expression.
3. Create an **ICE** table.
4. **Fill in** K's expression using the E row of the ICE table.
5. **Solve** for unknowns

I've written out the example from a YouTube video by Engineer4Free ([How to find equilibrium concentrations using an ICE table - YouTube](#))

Start with 1 M BrCl, which will undergo the reaction given.



Calculate the concentrations of all reaction components at equilibrium.

We already have the **balanced equation**. Let's start by writing an expression for **K**. Remember the equation for K:

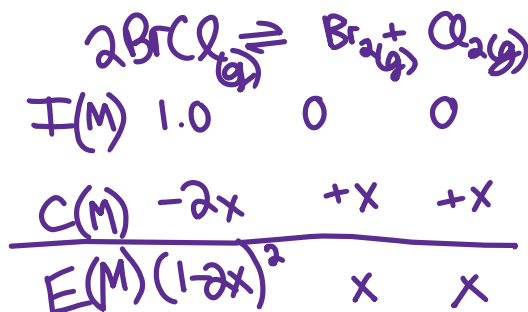
$$K_c = \frac{[C]^m * [D]^n}{[A]^j * [B]^k}$$

$$K_c = \frac{[\text{Br}_2][\text{Cl}_2]}{[\text{BrCl}]^2}$$

Next, set up an **ICE** table.

\*\*Put the initial concentration of BrCl (given) in the I row. Assume that the other two concentrations are 0 M.

\*\*Put the x under one of the two products – since their **coefficients are 1**, we won't have to deal with fractions. The change in concentration for BrCl is  $x \cdot (2/1)$ , and it is negative because BrCl is a reactant. The change in the other product concentration is calculated similarly.



The next step is to **plug the equilibrium concentrations (E row) into the  $K_c$  expression.**

$$K_c = \frac{[\text{Br}_2][\text{Cl}_2]}{[\text{BrCl}]^2} = \frac{x^2}{(1-2x)^2} = 5.81$$

Now, we can **solve for the unknown, x.** Multiply both sides by the denominator, combine like terms, and simplify.

$$x = 0.41 \text{ M}$$

X can be used to solve for all of the concentrations at equilibrium:

$$[\text{Br}_2] = x = 0.41 \text{ M}$$

$$[\text{Cl}_2] = x = 0.41 \text{ M}$$

$$[\text{BrCl}] = (1 - 2x)^2 = 0.18 \text{ M}$$

## Highlight 2: "X is Small"

For cases in which you are **solving for x**, as shown at the end of Highlight 1, and **the expression requires the quadratic formula.**

Here is an example from a YouTube video by Chantelle Anfuso ([The "x is small" Approximation - YouTube](#)):

- Consider the following reaction. If we start with 1.00 M of A, what are [A] and [B] at equilibrium?



$$K_c = \frac{[B]^2}{[A]} \quad 2.2 \times 10^{-3} = \frac{[B]^2}{[A]}$$

	[A] (M)	[B] (M)
Initial	1.00	0.0
Change	-x	+2x
Equilibrium	1.00 - x	2x

$$2.2 \times 10^{-3} = \frac{[2x]^2}{[1.00 - x]}$$

You can see that an ICE table has been set up, and the equilibrium row has been calculated. In the bottom right corner, the K expression has been filled in with the expressions from the equilibrium row. Notice: in contrast to the example from Highlight 1, the numerator is not  $x^2$ .

*(UNLESS! If you think that  $x$  will be less than  $0.05 \times \text{initial concentration}$ , so  $0.05 \times 1 \text{ M}$ , then you can remove the  $x$  term when it is being subtracted.  $x$  is small enough that, if subtracted from the initial concentration  $1 \text{ M}$ , it will not have a huge impact. But keep it on the numerator! A very small number, multiplied by itself, is an even smaller number. Once you've solved for  $x$ , check and see if it is actually less than  $0.05 \times 1 \text{ M}$ . If it is, it was all right to use the " $x$  is small" rule. If not, you need to go back and use the quadratic formula.)*

## Highlight 3: Significance of the Equilibrium Constant

All throughout CHE 1302, we come across equilibrium constant  $K_c$ . It is important to conceptually understand what  $K_c$  is and what it tells us about an equilibrium reaction! Recall -  $K_c$  is an equilibrium constant without any units, and it always follows this form:

$\frac{\text{product concentrations}}{\text{reactant concentrations}}$

- $K \gg 1$ : (e.g.  $K=110$ ), then the forward reaction essentially goes to completion – few reactants remain, much product formed
- $K > 1$ : (e.g.  $K = 10^{-14}$ ), then the forward reaction occurs to a slight extent – some reactants remain, product formed
- $K < 1$ : the reverse reaction is favored – higher concentration of reactants than products, little product formed
- $K \ll 1$ : the reverse reaction is very much favored – essentially no products, much reactant formed

NOTE: The size of  $K_c$  tells us what equilibrium will look like when we get there, but it says NOTHING about how quickly equilibrium is reached! How quickly equilibrium is reached is a different branch of chemistry – kinetics! Think back to rate laws.

## Check Your Learning

1. Determine the value of  $K_c$  for the reaction  $PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$  given the following information: At 525 K, a 20.82 g sample of  $PCl_5$  is placed in a 5.00 L container. At equilibrium there are 10.32 g of  $PCl_3$ .
2. For the reaction  $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$ ,  $K=107$ , is the forward or reverse reaction favored? Are the products or reactants favored?
3. Consider the following equilibrium reaction:  $4HCl(aq) + MnO_{2(s)} \rightleftharpoons MnCl_{2(aq)} + 2H_2O_l + Cl_{2(g)}$   $\Delta H = -272 \frac{kJ}{mol}$ . In which direction does the reaction occur if water is added?

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## Things You May Struggle With

1. Don't forget to balance the equation before using coefficients in a  $K$  expression!
2. Before doing ICE calculations: Compare  $Q$  and  $K$  so as to know which side of the equation has negative changes and which side of the equation has positive changes.

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That's all this week! Please reach out if you have any questions and don't forget to visit the Tutoring Center website for further information at [www.baylor.edu/tutoring](http://www.baylor.edu/tutoring). Answers to Check Your Learning are below.

1.  $K_c=0.0450$
2. Forward reaction; products
3. Left