Equilibrium Equations: Crash Course Chemistry #29

Crash Course: Chemistry
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=====Introduction (00:00)=====

Last week, I told you about equilibrium, which is proof if you ever needed proof that the universe is trying to mess with us, because even though we think and talk about chemical reactions as being a straightforward process, often times they're also going backwards.

At the same time, the reactants are combining to form products like nitrogen and hydrogen to make ammonia or hydrogen and fluorine to make hydrogen fluoride. The products are also breaking back apart into the reactants. So there's a sweet spot. One particular ratio of reactants to products where the products form at the same rate that they break down. When a reaction hits that spot, it's said to be in its equilibrium state. A bit of a kick in the head.

But of course you're not totally at chemistry's mercy; you can tinker with chemical equilibrium by altering the concentration of the substances and their temperature and if they're gases, the pressure on them. For example, adding pressure to the Haber Process that we use to create ammonia essentially shifts the position of equilibrium, so the result of the reaction is more ammonia being produced than nitrogen and hydrogen. But, wouldn't it be a lot more helpful if you knew how much pressure you'd have to add to produce the exact amount of ammonia you needed, or how much hydrogen fluoride it takes to refine a certain volume of gasoline. Math, of course, is the way to answer questions like these, so today I'm going to show you some simple, totally non-scary calculations that will help you get a handle on chemical equilibrium.

(Intro)

=====Calculating an Equilibrium Constant (01:21)=====

The first and most important thing you need to do equilibrium calculations is the equilibrium constant. This number is unique for every reaction and represents a molar ratio of products over reactants when a reaction is at equilibrium. Equilibrium constants are easy to set up but hard to explain, so let's start with an example using this obviously fake chemical equation.

The capital letters stand for the reactants and the products and the lowercase letters stand for their coefficients. The equilibrium constant, or $K_{eq,}$ is equal to the product of the molar concentration of the products divided by the product of the molar concentration of the reactants. Each concentration is raised to the power of its coefficient in the balanced equation.

Now this is important; we're using the coefficients as exponents here because we're multiplying all the products and all the reactants, not adding them like you would do in a balanced equation.

Generations of students have messed up test scores by getting this part wrong, but you will not do that. So again, it's the product of the products of the product of the reactants and the coefficients become exponents. Actually pretty simple. The square brackets in the formula are used by chemists to represent molar concentration, or molarity: moles of solute per liter of solution.

These equilibrium constant equations and the constants themselves are one of the few places where you don't need to write them with every number. Just remember to convert everything into molarities before plugging numbers into the equation.

One last thing before we do a calculation: as we learned in the last episode, a change in temperature changes the position of equilibrium. Therefore, the equilibrium constant is only true for a specific temperature. Constants are normally calculated at 25 degrees Celsius, which is close enough for most situations, but the temperature should always be mentioned along with the constant.

=====Calculating Conditions of Reactions (03:00)=====

Fortunately for us, chemists have already figured out the equilibrium constants for most common reactions. Carbonic acid, for instance, which is basically just carbon dioxide dissolved in water, dissociates to form carbonate ions and hydrogen ions. You'll see this reaction again later when we talk in depth about carbon and the planet's carbon cycles. More importantly, right now, the reaction is perfectly reversible with an equilibrium constant of 1.66×10^{-17}. For this reaction, K_{eq} equals the product of the molar concentration of the carbonate ion and the molar concentration of the hydrogen ion, which is squared because hydrogen's coefficient in the balanced equation is two, all divided by the molar concentration of carbonic acid.

Now, let's not be boring and throw a bunch of numbers around. What you need to see is that the quotient on the right must always yield the same number because it's a constant. So, if the amount of CO_2 in the atmosphere increases leading to more H_2CO_3 in the earth's water, the concentration of carbonate ions and or hydrogen ions must also increase so the total will match the K_{eq}. You should also notice that any change to the concentration of hydrogen ions will have a huge effect on the denominator because its square.

So an increase in hydrogen ions, like if the water were somehow acidified by an outside source, would require a huge decrease in carbonate ions or a huge increase in carbonic acid—carbon dioxide pulled in from the atmosphere to satisfy the equilibrium condition again.

=====RICE Tables (04:21)=====

We can attack this problem from another direction, too, depending on what information we have to begin with. Chemists often know not only the equilibrium constant for a reaction, but also how much of each reactant is available. And all they need to figure out is exactly how much of each substance will be present at equilibrium.

This type of calculation is easiest using a format called a RICE table. RICE stands for Reaction, Initial, Change, and Equilibrium.

On the R line at the top of the table, we write the chemical equation of reaction, leaving space between each part so we'll have room to add more information below.

On the I line we write the initial concentrations of each substance. Some of those will almost always be zero, since products generally aren't present until the reaction begins.

The C line is where we map out how much of each substance will change during the reaction. We often don't know exactly how much this is until we do the math, so we start out with x, where the amount is unknown.

The E line is where we put the final result: how much of each substance will be present at equilibrium. Since the final amount is just the initial amount plus any changes that have occurred, this line is the sum of the initial line and the change line.

Let's do this for hydrogen fluoride, or HF, an integral part of the process of refining gasoline. It can be formed in the gaseous state by an equilibrium reaction between hydrogen gas and fluorine gas.

Start by writing the balanced equation. For our initial concentrations, let's use 3.00 mol of H_2 and 6.00 mol of F_2 in a 3.00 liter container at a certain temperature. That makes the initial concentration of H_2 3 moles in 3 liters, or a 1.00 molar solution. Similarly, the initial concentration of the F_2 is 6 moles per 3 liters, or 2 molar. No HF has formed yet, so it's initial concentration is 0.

So what we're trying to figure out is how much of the hydrogen and the fluorine will react to form hydrogen fluoride under these conditions. So we'll call the change 'x' for now.

Since 1 mole of H2 and 1 mole of F_2, combine to form 2 moles of HF. We can say that the H_2 and F_2 will each lose 'x' moles per liter while the HF gains 2x moles per liter. That's on the change line of your table.

So that leaves the equilibrium line where you write your totals. For H_2 we have a total of 1.00 - x molar. For F_2 the total is 2.00 - x. For the HF the total is 2x molar.

=====Quadratic Equations (06:35)=====

Now we apply these figures to our formula for the equilibrium constant. Based on the table in the back of my textbook, the Keq for this reaction is 115 at the given temperature. We plug in the numbers from our RICE table and solve for x. And we end up with this. Which you probably recognize as a quadratic equation.

Just so you know that doesn't happen with every equilibrium calculation, but it's an extremely common result. To solve the quadratic equation we have to use the quadratic formula. And to do that we have to think of the coefficients of our equation as a, b, and c in that order. Then we plug them in the corresponding positions in this formula: -b +/- the square root of b2 - 4ac, all divided by 2a.

Once we finally get all the numbers in the right places, it becomes a matter of grinding through some basic algebra. And because of the way the quadratic formula works, we'll get two possible answers.

To figure out which one is the right one, think back to the beginning of the problem. I said the initial concentration of the H_2 was 1 molar, and the initial concentration of the F_2 was 2 molar. And x is the amount that each one lost, right?

Well neither the H_2 nor the F_2 could possibly lose 2.13 moles per liter when they both started with less than that already. So clearly, the correct answer has to be 0.968 molar.

Using that then, we can calculate the actual equilibrium amounts for each substance. At equilibrium, under these specific conditions the concentration of HF is 2x, or 2 x 0.968, which equals 1.94 molar. The concentration of hydrogen will be 1 - 0.968 or 0.032 molar. And the concentration of fluorine will be 2 - 0.968 or 1.03 molar.

As we learned last week, those values can be shifted. For instance, to maximize the hydrogen fluoride production by adding or removing some of the substance. Or by changing the pressure or temperature on the system.

That's why these calculations are so valuable to scientists. Just a little bit of algebra allows us to maximize the benefit that we receive from the things that nature is already doing anyway. If the universe really is try to take advantage of us, we've certainly figured out how to take advantage right back.

So the next time you're wondering why you learned this stuff in math class, now you know.

=====Summary & Credits (08:41)=====

Thanks for watching this episode of CrashCourse Chemistry. Today you've learned some mathematical tools to help us make more efficient use of equilibrium reactions in real life. You've learned how to calculate an equilibrium constant, K_{eq}. You've learned how to calculate the equilibrium conditions of reactions just from knowing their initial conditions. And you've learned that a RICE table isn't just a place where you eat sushi. And you may even have learned a little bit about the quadratic equation.

This episode was written by Edi González. The script was edited by Blake de Pastino. And our chemistry consultant was Dr. Heiko Langner. It was filmed, edited, and directed by Nicholas Jenkins. The script supervisor was Katherine Green. The sound designer is Michael Aranda. And our graphics team is Thought Café.