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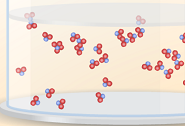
**Student Exploration: Equilibrium and Concentration**

**Vocabulary:** chemical equilibrium, concentration, equilibrium, equilibrium constant, reaction quotient, reversible reaction

**Prior Knowledge Questions** (Do these BEFORE using the Gizmo.)

Gary has $5,000 in his bank account and earns a modest salary. Every month he pays for rent, food, utilities, and entertainment.

1. How will Gary’s account change if he saves more than he spends? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
2. How will Gary’s account change if he spends more than he saves? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
3. What happens if Gary spends exactly as much as he saves? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Gizmo Warm-up**

If Gary spends exactly as much as he earns, his savings will be in **equilibrium**. Equilibrium occurs when two opposing processes occur at the same rate, leading to no net change. In the *Equilibrium and Concentration* Gizmo, you will investigate how equilibrium can occur in chemical reactions.

To begin, check that **Reaction 1** is selected. Set **Moles NO2** to 8 and **Moles N2O4** to 0.

1. Click **Play** (Play) and observe the colliding molecules. What do you notice? \_\_\_\_\_\_\_\_\_\_\_\_\_\_

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In the Gizmo, a blue flash appears every time two reactants combine to form a product. A red flash appears every time a product dissociates into reactants.

1. Click **Reset** (Reset), and set **Moles NO2** to 0 and **Moles N2O4** to 8. Click **Play**.

What do you notice now? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. When a reaction can proceed in either direction, it is a **reversible reaction**. Based on what you have observed, is the synthesis of NO2 into N2O4 a reversible reaction? Explain.

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| **Activity A:**  **Reversible reactions** | Get the Gizmo ready:   * Click **Reset**. **Reaction 1** should be selected. * Set **Moles NO2** to 8 and **Moles N2O4** to 0. * Move the **Sim. speed** slider all the way to the right. |  |

**Question: What are the characteristics of reversible reactions?**

1. Predict: Suppose you began with 8 moles of NO2 in the chamber. What do you think will happen if you let the reaction go for a long time? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

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1. Test: Click **Play**. Select the BAR CHARTtab and check that **Moles** is selected. Observe the bar chart for about 30 seconds. As time goes by, what do you notice about the bars representing moles NO2 and moles N2O4?

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1. Observe: Click **Pause** (Pause). Select the GRAPH tab. Click the (–) zoom control on the horizontal axis until you can see the whole graph. What do you notice?

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This situation, in which the overall amounts of reactants and products does not change significantly over time, is called a **chemical equilibrium**.

1. Record: On the BAR CHARTtab, turn on **Show data values**. How many moles of NO2 and N2O4 are there right now? Moles NO2 \_\_\_\_\_\_\_\_\_\_ Moles N2O4 \_\_\_\_\_\_\_\_\_\_
2. Calculate: Suppose all the NO2 molecules were synthesized into N2O4. Given the equation 2NO2 ⇄ N2O4, how many moles of N2O4 would be produced? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
3. Experiment: Click **Reset**. On the INITIAL SETTINGStab, set **Moles NO2** to 0 and **Moles N2O4** to 4. Click **Play**. Click **Pause** when the bars of the bar chartstop moving very much.
4. List the current amounts of each substance: Moles NO2 \_\_\_\_\_\_ Moles N2O4 \_\_\_\_\_\_
5. How do these results compare to starting with 8 moles of NO2? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

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**(Activity A continued on next page)Activity A (continued from previous page)**

1. Summarize: In each trial, you started with the same amounts of nitrogen and oxygen. In this situation, did the equilibrium amounts change depending on the direction of the reaction?

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1. Set up the Gizmo: Click **Reset** and select the EXPERIMENTtab on the left. On the INITIAL SETTINGStab on the right, select **Reaction 2**. Set **Moles NO** to 5, **Moles NO2** to 5, and **Moles N2O3** to 0. What are the reactants and product of this reaction?

Reactants: \_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_ Product: \_\_\_\_\_\_\_\_\_\_

(Note: In this reaction, some of the NO2 reactants combine to form N2O4, as in reaction 1.)

1. Observe: Recall that a blue flash appears every time two reactants combine to form a product. A red flash appears every time a product dissociates into reactants. Click **Play**.
2. At first, do you notice more blue flashes or red flashes? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
3. What do you notice about the frequency of blue and red flashes as time goes by?

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1. Click **Reset**. This time, start the experiment with 0 moles of NO and NO2 and 5 moles of N2O3. Click **Play**. What do you notice about the red and blue flashes now?

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1. Explain: Think about how the numbers of blue and red flashes reflect the rates of the forward (reactants 🡪 products) and reverse (products 🡪 reactants) reactions.
2. What happens to the rate of the forward reaction as the reactants are consumed?

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1. What happens to the rate of the reverse reaction as the products are produced?

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1. Why do reversible reactions *always* result in chemical equilibria? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

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| **Activity B:**  **The equilibrium constant** | Get the Gizmo ready:   * Click **Reset**. Select **Reaction 1**. * Set **Moles NO2** to 2 and **Moles N2O4** to 7. | 1046SE3 |

**Introduction:** When investigating the rates of reactions, it often is useful to consider the **concentrations** of reactants rather than the total number of moles. Concentrations are often expressed in moles per liter, or mol/L. Brackets are used to signify concentration. For example, “[H2] = 5.0 M” means the concentration of hydrogen gas in a chamber is 5.0 moles per liter.

**Question: What are the characteristics of reactions in equilibrium?**

1. Record: On the BAR CHARTtab, select **Concentration**. Check that **Show data values** is on. If necessary, use the arrows to adjust the scale of the chart.
2. What are the current concentrations of each compound?

[NO2] \_\_\_\_\_\_\_\_\_\_ [N2O4] \_\_\_\_\_\_\_\_\_\_

1. Click **Play** and wait for equilibrium to become established. Click **Pause**. What are the approximate equilibrium concentrations?

[NO2] \_\_\_\_\_\_\_\_\_\_ [N2O4] \_\_\_\_\_\_\_\_\_\_

1. Calculate: The value *Kc* represents the ratio of products to reactants in a reaction at equilibrium. The greater the amount of products relative to reactants, the higher the resulting value of *Kc*. For a general reaction between gases: *a*A(*g*) + *b*B(*g*) ⇌ cC(*g*) + *d*D(*g*), *Kc* is calculated as follows:



For the current reaction, 2NO2 ⇌ N2O4, we have:



Based on the current concentrations of NO2 and N2O4, what is *Kc*? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Show your work here:

**(Activity B continued on next page)Activity B (continued from previous page)**

1. Gather data: Experiment with a variety of initial concentrations of NO2 and N2O4. For each set of initial concentrations, use the Gizmo to determine the equilibrium concentrations of each substance. In the last column, find *Kc* for that trial. Run three trials for each set of initial conditions.

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| **Initial  [NO2]** | **Initial  [N2O4]** | **Equilibrium [NO2]** | **Equilibrium [N2O4]** | ***Kc*** |

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1. Calculate: Find the average value of *Kc* for each set of three trials.

Trials 1-3: \_\_\_\_\_\_\_\_\_\_ Trials 4-6: \_\_\_\_\_\_\_\_\_\_ Trials 7-9: \_\_\_\_\_\_\_\_\_\_

1. Analyze: What do you notice about the values of *Kc*? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

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In general, the value of *Kc* will be constant for a given reaction at a constant temperature, no matter the starting concentrations. That is why *Kc* is known as the **equilibrium constant**. In this Gizmo, the values of *Kc* will vary somewhat because there is a very limited number of molecules in the chamber.

1. On your own: Use the Gizmo to find *Kc* for **Reaction 4**: H2 + I2 ⇌ 2HI. Collect data at least 10 times and average your results to get the best approximation of *Kc*. Show your data and work on a separate sheet of paper.

(Hint: Because of the coefficient “2” in front of HI, you will have to square the concentration of HI to find *Kc*.)

*Kc* = \_\_\_\_\_\_\_\_\_\_

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| **Activity C:**  **Reaction direction** | Get the Gizmo ready:   * Click **Reset**. Check that **Reaction 4** is selected. * Set **Moles H2** to 5, **Moles I2** to 5, and **Moles HI**  to 3. |  |

**Introduction:** For a reversible reaction with equilibrium constant *Kc*, it often is useful to know in which direction the reaction will proceed given the starting amounts of reactants A and B and products C and D. This is done by calculating the **reaction quotient**, *Qc*:



**Question: How can you predict the direction of a reversible reaction?**

1. List: Select the BAR CHARTtab. What are the initial concentrations of each substance?

[H2] \_\_\_\_\_\_\_ [I2] \_\_\_\_\_\_\_ [HI] \_\_\_\_\_\_\_

1. Calculate: Use the equation above to find *Qc* for the current reaction.
2. What is the current value of *Qc*? \_\_\_\_\_\_\_\_\_\_
3. In activity B, what value of *Kc* did you arrive at for this reaction? \_\_\_\_\_\_\_\_\_\_
4. How does *Qc* compare to *Kc*? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
5. Analyze: Recall that *Qc* is equal to the ratio of product concentrations to reactant concentrations.
6. If there is an excess of products, will *Qc* be greater than or less than *Kc*? \_\_\_\_\_\_\_\_\_\_
7. If there is an excess of reactants, will *Qc* be greater than or less than *Kc*? \_\_\_\_\_\_\_\_\_
8. In the current situation, is there an excess of products or reactants? \_\_\_\_\_\_\_\_\_\_\_\_\_

Explain: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. When the reaction begins, do you expect [HI] to increase or decrease? \_\_\_\_\_\_\_\_\_\_\_

Explain: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Test: Click **Play**. What happens to [HI]? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

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| **Extension:**  **Equilibrium calculations** | Get the Gizmo ready:   * Click **Reset**. Select **Reaction 1**. * Set **Moles NO2** to 0 and **Moles N2O4** to 6. |  |

**Goal: Given *Kc* and initial concentrations, calculate equilibrium concentrations.**

1. List: Select the BAR CHART. What is the initial concentration N2O4? [N2O4]*initial* = \_\_\_\_\_\_\_
2. Experiment: Click **Play** and wait for a few seconds. Click **Pause** before equilibrium is reached.
3. What is the current concentration of N2O4? [N2O4] = \_\_\_\_\_\_\_
4. How much has the concentration of N2O4 gone down? \_\_\_\_\_\_\_
5. What is the current concentration of NO2? [NO2] = \_\_\_\_\_\_\_
6. In general, if [N2O4] is reduced by *x*, how much does [NO2] increase? \_\_\_\_\_\_\_\_\_\_\_\_

This result may be surprising. It is true because at constant pressure, the overall density of particles in the container remains constant. So, if the concentration of one substance is reduced by *x*, the concentration of the other substance increases by *x*.

1. Manipulate: Begin with the general equation for *Kc*: .
2. What is the equation for *Kc* for the reaction 2NO2 ⇌ N2O4? *Kc* =
3. In this experiment, the initial concentration of NO2 is zero. If the concentration of N2O4 is reduced by *x* at equilibrium, the equilibrium concentration of NO2 is equal to *x*. Substitute the following values into the equation you wrote in step A:

[N2O4] = ([N2O4]*initial* – *x*) [NO2] = *x*

*Kc* =

1. In activity A, you discovered that *Kc* for this reaction was close to 0.042. Substitute this value and the initial concentration of N2O4 into your equation.

=

1. Rearrange the terms of your equation to form a quadratic equation in the form   
   *ax2* + *bx* + *c* = 0.

= 0

**(Extension continued on next page)Extension (continued from previous page)**

1. Solve: Because the equation is in the form *ax2* + *bx* + *c* = 0, you can use the quadratic formula (shown below) to solve for *x*. Ignore negative solutions because the concentrations cannot be negative. Show your work.



1. Predict: Based on the value for *x*, what do you expect the equilibrium concentrations of NO2 and N2O4 to be?

[NO2] \_\_\_\_\_\_\_\_\_\_ [N2O4] \_\_\_\_\_\_\_\_\_\_

Check your work by solving for *Kc* using  *Kc* = \_\_\_\_\_\_\_\_\_\_

If you don’t get the correct value of *Kc*, recheck your work.

1. Test: Click **Play** and wait for equilibrium to be established. What are the actual equilibrium values of each substance?

[NO2] \_\_\_\_\_\_\_\_\_\_ [N2O4] \_\_\_\_\_\_\_\_\_\_

How close were these results to your predicted results? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

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1. Challenge: Suppose you begin with 6 moles of NO2 and 5 moles of N2O4. Assuming a value for *Kc* of 0.042, predict the equilibrium concentrations of NO2 and N2O4. (Use the Gizmo to determine the initial concentrations.) Show your work on a separate sheet of paper. After you have made your predictions, click **Play** and record the experimental results.

Predicted: [NO2] \_\_\_\_\_\_\_ [N2O4] \_\_\_\_\_\_\_

Experimental: [NO2] \_\_\_\_\_\_\_ [N2O4] \_\_\_\_\_\_\_