Task 3

Due: 11:59pm on Friday, April 27, 2018

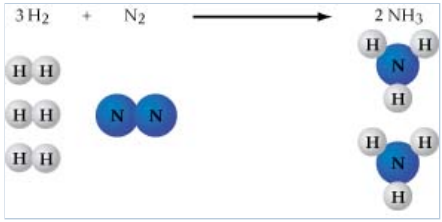
To understand how points are awarded, read the Grading Policy for this assignment.

Balancing Chemical Equations

Learning Goal:

To learn to balance chemical equations by inspection.

According to the law of conservation of mass, matter cannot be created or destroyed. Therefore, a chemical equation must show the same number of each kind of atom in the reactants as it does in the products. As shown in the figure, the balanced equation $3H_2 + N_2 \rightarrow 2NH_3$ has 6 atoms of hydrogen and 2 atoms of nitrogen on each side of the arrow.



One is often presented with unbalanced chemical equations for which one must supply the coefficients.

The following equation is not balanced:

$$CaSO_4 + AlCl_3 \rightarrow CaCl_2 + Al_2(SO_4)_3$$

Part A

Notice that "SO_4" appears in two different places in this chemical equation. SO_4^{2-} is a *polyatomic ion* called "sulfate." What number should be placed in front of $CaSO_4$ to give the same total number of sulfate ions on each side of the equation?

$$?CaSO_4 + AlCl_3 \rightarrow CaCl_2 + Al_2(SO_4)_3$$

Express your answer numerically as an integer.

Hint 1. Confused by parentheses?

Parentheses surrounding a set of elements means that the subscript applies to all the atoms within the parentheses. For example: the compound $Al_2(SO_4)_3$ contains two aluminum atoms and three sulfate ions (12 oxygen atoms, 3 sulfur atoms).

ANSWER:

? = 3

Correct

Part B

Now that we have put a coefficient of 3 in front of $CaSO_4$, what coefficient should go in front of $CaCl_2$ to balance calcium (Ca)?

$$3CaSO_4 + AlCl_3 \rightarrow ?CaCl_2 + Al_2(SO_4)_3$$

Express your answer numerically as an integer.

Hint 1. Count the number of calcium atoms on the left

How many calcium atoms currently appear on the left side of this equation?

$$3CaSO_4 + AlCl_3 \rightarrow CaCl_2 + Al_2(SO_4)_3$$

Express your answer numerically as an integer.

ANSWER:

3

ANSWER:

? = 3

Correct

By placing a coefficient of 3 in front of calcium chloride, we have indicated three calcium atoms on the right to match the three calcium ions on the left:

$$3CaSO_4 + AlCl_3 \rightarrow 3CaCl_2 + Al_2(SO_4)_3$$

Typesetting math: 68%

Part C

Now that we have put coefficients of 3 in front of $CaSO_4$ and $CaCl_2$, what coefficient should go in front of $AlCl_3$ to balance both the Cl atoms and the Al atoms?

$$3CaSO_4 + ?AlCl_3 \rightarrow 3CaCl_2 + Al_2(SO_4)_3$$

Express your answer numerically as an integer.

ANSWER:

? = 2

Correct

There are now 6 chlorine atoms, and 2 aluminum atoms on each side of the equation. The final balanced equation looks like this:

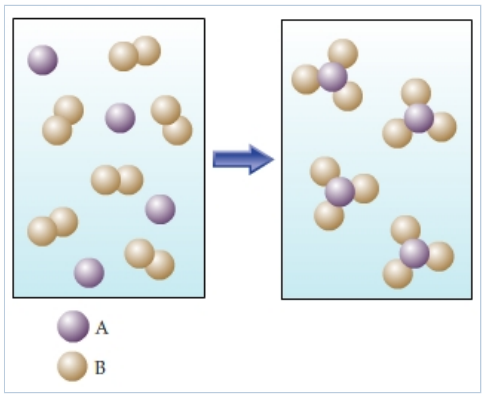
$$3CaSO_4 + 2AlCl_3 \rightarrow 3CaCl_2 + Al_2(SO_4)_3$$

Chemical Equations

When a chemical reaction occurs, atoms rearrange to form new compounds, but no new atoms are created nor are any destroyed. This concept is called *conservation of mass*. Mass conservation can be seen in a balanced chemical equation, where the numbers of each kind of atom are the same on both sides of the reaction arrow.

Part A

The figure shows the reaction of element A (lavender spheres) with element B (tan spheres). Write the balanced chemical equation for this reaction in terms of A and B.



Express your answer as a chemical equation.

Hint 1. Identify the product

What is the chemical formula of the product?

Express your answer as a chemical formula.

ANSWER:

AB_3

Hint 2. Identify the reactants

What are the chemical formulas of the reactants?

Express your answers as chemical formulas separated by a comma.

ANSWER:

A,B_2

Hint 3. Determine the coefficients

Here is the unbalanced chemical equation:

$$?\,A + ?\,B_2 \rightarrow ?\,AB_3$$

Determine the coefficients for the balanced equation.

Enter three integers separated by commas.

ANSWER:

2,3,2

ANSWER:

$2A + 3B_2 \rightarrow 2AB_3$

Correct

Typesetting math: 68%

Balancing equations with polyatomic ions

When balancing chemical equations with polyatomic ions, it may help to balance the ion as a whole. For example, the ion PO_4^{3-} can be treated as a unit and balanced in both the reactants and the product, or you can break it down and balance each atom separatly.

Part B

Phosphoric acid reacts with magnesium hydroxide to produce magnesium phosphate and water via the following reaction:

$$?\,H_3PO_4 + ?\,Mg(OH)_2 \rightarrow ?\,Mg_3(PO_4)_2 + ?\,H_2O$$

Balance the equation, then enter the coefficients, in order, in the answer box.

Enter four integers separated by commas.

Hint 1. Balance magnesium

What coefficient should be placed in front of $Mg(OH)_2$ to balance the magnesium atoms?

Express your answer numerically as an integer.

ANSWER:

3

Hint 2. Balance phosphorus

Consider this partially balanced equation:

$$?\,H_3PO_4 + 3Mg(OH)_2 \rightarrow Mg_3(PO_4)_2 + ?\,H_2O$$

What coefficient should be placed in front of H_3PO_4 to balance the phosphorus atoms?

Express your answer numerically as an integer.

ANSWER:

2

Hint 3. Balance hydrogen

Consider this partially balanced equation:

$$2H_3PO_4 + 3Mg(OH)_2 \rightarrow Mg_3(PO_4)_2 + ?\,H_2O$$

What coefficient should be placed in front of H_2O to balance the hydrogen atoms?

Express your answer numerically as an integer.

Hint 1. Count the number of H atoms shown on the left side

How many H atoms are currently shown on the left side of the equation?

Express your answer numerically as an integer.

Hint 1. Count the H atoms in $2H_3PO_4$

How many H atoms are in $2H_3PO_4$?

Express your answer numerically as an integer.

ANSWER:

6

Hint 2. Count the H atoms in $3Mg(OH)_2$

How many H atoms are in $3Mg(OH)_2$?

Express your answer numerically as an integer.

ANSWER:

6

ANSWER:

12

ANSWER:

6

ANSWER:

2,3,1,6

Correct

The balanced equation is written as follows:

$$2H_3PO_4 + 3Mg(OH)_2 \rightarrow Mg_3(PO_4)_2 + 6H_2O$$

Notice that the equation shows a total of 12 hydrogen atoms, 2 phosphorus atoms, 3 magnesium atoms, and 14 oxygen atoms on each side.

Part C

You carefully weigh out 16.00 g of $CaCO_3$ powder and add it to 64.80 g of HCl solution. You notice bubbles as a reaction takes place. You then weigh the resulting solution and find that it has a mass of 74.24 g . The relevant equation is

$$CaCO_3\,(s) + 2HCl\,(aq) \rightarrow H_2O\,(l) + CO_2\,(g) + CaCl_2\,(aq)$$

Assuming no other reactions take place, what mass of CO_2 was produced in this reaction?

Express your answer to three significant figures and include the appropriate units.

Hint 1. How to approach the problem

Matter is neither created nor destroyed in a chemical reaction. Therefore, the total mass of the reactants must be equal to the total mass of the products. First, find the initial mass by adding together the masses of the reactants. The final mass (74.24 g) is less than the initial mass because CO_2 gas escaped during the reaction. The mass of CO_2 that escaped is equal to the difference between the initial and final masses.

Hint 2. Determine the initial mass

What was the combined mass of the HCl solution and the $CaCO_3$, before the reaction started?

Express your answer to four significant figures and include the appropriate units.

Hint 1. How to find the initial mass

Find the initial mass by adding together the masses of the reactants, HCl and $CaCO_3$, which were weighed out before the reaction.

ANSWER:

80.80 g

Correct

ANSWER:

6.56 g

Correct

Calculations Using the Mole

Learning Goal:

To learn how to convert grams to moles and to use the mole to find the number of atoms in a sample.

The *mole* (abbreviated mol) is a counting unit used to simplify calculations that would otherwise involve very large numbers. The mole is equivalent to the number of carbon atoms in exactly 12 g of isotopically pure ^{12}C, or 6.02×10^{23}. This number is known as *Avogadro's number* in honor of Amedeo Avogadro.

Avogadro's number can be used as a conversion factor between moles and atoms as shown here:

$$\frac{6.02\times10^{23}\text{ atoms}}{1\text{ mole of atoms}}$$

The *molar mass* of a substance is the mass of one mole of a substance and is written in units of grams per mole. The molar mass of an atom is equivalent to its atomic mass whereas the molar mass of a substance is equivalent to its formula weight.

Part A

How many moles of atoms are in 5.00 g of ^{13}C?

Express your answer numerically in moles.

Hint 1. Determine the molar mass of carbon-13

What is the molar mass of ^{13}C?

Express your answer numerically in grams per mole.

ANSWER:

13.0 $\mathrm{g/mol}$

Hint 2. Identify the operation that allows units to cancel properly

Given x g of a substance that has a molar mass of y $\mathrm{g/mol}$, which operation will give an answer with units of moles?

ANSWER:

- $x \cdot y$
- $\frac{x}{y}$
- $\frac{y}{x}$

ANSWER:

0.385 $\mathrm{mol\ ^{13}C}$

Correct

Part B

Based on your answer in Part A, calculate the number of atoms in this amount of $^{13}\mathrm{C}$?

Express your answer numerically in atoms of carbon.

Hint 1. How to approach the problem

Use Avogadro's number to convert moles of $^{13}\mathrm{C}$ to atoms of carbon. Recall that Avogadro's number is 6.02×10^{23}. In other words, each mole of atoms contains 6.02×10^{23} atoms.

ANSWER:

2.32×10^{23} $^{13}\mathrm{C}$ atoms

Correct

Part C

Based on your answer in Part B, how many electrons are in this amount of $^{13}\mathrm{C}$?

Express your answer numerically in electrons.

Hint 1. How to approach the problem

In a neutral atom, the number of protons is equal to the number of electrons. The number of protons in an atom can be determined by looking up the atomic number of the element in the periodic table.

Once you know the number of electrons per atom, multiply by the total number of atoms, 2.32×10^{23}.

Hint 2. Determine the number of electrons in one atom of carbon-13

How many electrons are in one atom of $^{13}\mathrm{C}$?

Express your answer numerically as an integer.

ANSWER:

6 electrons

ANSWER:

1.39×10^{24} electrons

Correct

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Part D

Based on your answer in Part B, how many neutrons are in this amount of ^{13}C?

Express your answer numerically in neutrons.

Hint 1. How to approach the problem

The mass number of an atom is the sum of its protons and neutrons. The number of protons in an atom can be determined by looking up the atomic number of the element in the periodic table.

Once you know the number of neutrons per atom, multiply by the total number of atoms, 2.32×10^{23}.

Hint 2. Determine the number of neutrons in one atom of carbon-13

How many neutrons are in one atom of ^{13}C?

Express your answer numerically as an integer.

ANSWER:

7 neutrons

ANSWER:

1.62×10^{24} neutrons

Correct

± Empirical Formula Procedure

Learning Goal:

To understand what an empirical formula is and to learn the procedure for finding empirical formulas.

An empirical formula expresses the simplest ratio of the elements involved. The compound P_4O_{10} has a P:O ratio of 4:10. However, this ratio can be simplified to 2:5. Therefore the empirical formula for P_4O_{10} is P_2O_5.

Another example is hydrogen peroxide, which has the formula H_2O_2 and an empirical formula of HO.

One way that chemists analyze new or unknown compounds is to determine the percentage composition of elements experimentally, which yields the empirical formula.

Part A

What is the empirical formula for the compound P_4O_6?

Express your answer as a chemical formula.

ANSWER:

P_2O_3

Correct

Information for Parts B, C, and D

A compound is 80.0% carbon and 20.0% hydrogen by mass. Assume a 100.-g sample of this compound.

Part B

How many grams of each element are in this sample?

Enter the number of grams of carbon followed by the number of grams of hydrogen, separated by a comma (e.g., `30.0,70.0`).

ANSWER:

mass of C, mass of H = 80.0,20.0 g

Correct

Part C

How many moles of each element are in this sample?

Enter the number of moles of carbon followed by the number of moles of hydrogen, separated by a comma.

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Hint 1. How to convert grams to moles

To convert from 80.0 grams of carbon to moles of carbon, divide 80.0 by the molar mass of carbon.
To convert 20.0 grams of hydrogen to moles of hydrogen, divide 20.0 by the molar mass of hydrogen.

ANSWER:

moles of C, moles of H = 6.66,19.8 mol

Correct

Part D

Based on the mole ratio you determined in Part C, what is the empirical formula of this compound?

Express your answer as a chemical formula.

Hint 1. How to approach the problem

Your answer from Part C may be expressed as a mole ratio of $\mathrm{C:H} = 6.66:19.8$. However, these numbers are not appropriate to use as subscripts in the empirical formula because they are not integers.

Multiplying or dividing each value in the mole ratio by the same number will not change the ratio.

Hint 2. Find the integer mole ratio

What is the simplest integer mole ratio of carbon to hydrogen, $\mathrm{C:H}$, for this compound?

Enter the number of moles of C followed by the number of moles of H, separated by a colon (e.g., 5:6).

Hint 1. Ratio example

Just as the ratio 1:2 is equal to the ratio 2:4, it is also equal to the ratio 0.5:1. The second number is always twice the first number: $2/1 = 2$, $4/2 = 2$, and $1/0.5 = 2$.

ANSWER:

$\mathrm{C:H}$ = 1:3

ANSWER:

$\mathrm{CH_3}$

Correct

± Percent Composition

Percent composition refers to the mass percent of each element in a compound:

$$\text{mass percent} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100\%$$

For example, the percent composition of water, $\mathrm{H_2O}$, is 11.2% hydrogen and 88.8% oxygen. Therefore, a 100-g sample of water contains 11.2 g of hydrogen atoms and 88.8 g of oxygen atoms.

The periodic table will be useful when doing this problem. You can access a periodic table by clicking the "Tools" link in the upper right corner of this page.

Part A

A hydrocarbon is a compound that contains mostly carbon and hydrogen. Calculate the percent composition (by mass) of the following hydrocarbon: $\mathrm{C_{10}H_{22}}$.

Enter the percentages of carbon and hydrogen numerically, separated by commas.

Hint 1. How to approach the problem

Start by assuming that you have exactly 1 mol of the compound. Then calculate the mass of carbon and the mass of hydrogen in this sample. Finally, make a percentage out of the mass of each element and the mass of the compound.

Hint 2. Calculate the molar mass of the compound

What is the molar mass of $\mathrm{C_{10}H_{22}}$?

Express your answer numerically in grams per mole using four significant figures.

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Hint 1. Definition of molar mass

The molar mass of $C_{10}H_{22}$ is the sum of the masses of 10 mol of C and 22 mol of H.

ANSWER:

142.3 $\mathrm{g/mol}$

Hint 3. Calculate the mass of carbon

Calculate the mass of carbon in exactly 1 mol of $C_{10}H_{22}$.

Express your answer numerically in grams using four significant figures.

Hint 1. How to approach the problem

There are 10 mol C in 1 mol $C_{10}H_{22}$. Each mole of C has a mass of 12.01 g.

ANSWER:

120.1 g

Hint 4. Calculate the mass of hydrogen

Calculate the mass of hydrogen in exactly 1 mol of $C_{10}H_{22}$.

Express your answer numerically in grams using four significant figures.

ANSWER:

22.2 g

ANSWER:

carbon, hydrogen = 84.4,15.6 $\%$

Correct

Notice that the sum of the percentages is 100$\%$.

Part B

A certain metal hydroxide, $\mathrm{M(OH)_2}$, contains 32.8$\%$ oxygen by mass. What is the identity of the metal M?

Enter the full name of the element.

Hint 1. How to approach the problem

Assume that you have exactly 1 mol of $\mathrm{M(OH)_2}$ and calculate the mass of oxygen present. Then use the mass percent formula to calculate the molar mass of the compound. Once you know the molar mass of the compound, you can figure out the molar mass of the metal by subtracting the mass of the hydroxide ions.

Hint 2. Calculate the molar mass of the metal hydroxide

Calculate the molar mass of $\mathrm{M(OH)_2}$ using the fact that it contains 32.8$\%$ oxygen by mass.

Express your answer in grams per mole using four significant figures.

Hint 1. How to rearrange the mass percent formula

If you assume 1 mol of compound, the formula from the introduction becomes

$$\text{mass percent of O} = \frac{\text{mass of O in 1 mol of compound}}{\text{molar mass of compound}} \times 100\%$$

This can be rearranged to solve for the molar mass of the compound:

$$\text{molar mass of compound} = \frac{\text{mass of O in 1 mol of compound}}{\text{mass percent of O}} \times 100\%$$

Hint 2. Calculate the mass of oxygen in one mole of the compound

How many grams of oxygen are in 1 mol of $\mathrm{M(OH)_2}$?

Express your answer numerically in grams.

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Hint 1. How to approach the problem

First determine the number of moles of O in 1 mol of $\mathrm{M(OH)_2}$. Use the periodic table to determine the mass of each mole of O (the molar mass).

Hint 2. Determine the number of moles of oxygen

How many moles of oxygen are in 1 mol of $\mathrm{M(OH)_2}$?

Express your answer as an integer.

ANSWER:

2

Hint 3. Calculate the mass of one mole of oxygen

What is the mass of 1 mol of O?

Express your answer in grams to four significant figures.

ANSWER:

16.00 g

ANSWER:

oxygen mass = 32.0 g

ANSWER:

molar mass = 97.56 $\mathrm{g/mol}$

Hint 3. Calculate the molar mass of the metal

A certain metal hydroxide, $\mathrm{M(OH)_2}$, has a molar mass of 97.6 $\mathrm{g/mol}$. Find the molar mass of the metal M.

Express your answer in grams per mole using four significant figures.

Hint 1. How to approach the problem

You have already calculated the molar mass of the compound $\mathrm{M(OH)_2}$. Now determine the molar mass of $\mathrm{(OH)_2}$. The difference between the two numbers represents the molar mass of M:

$$\text{molar mass of M} = [\text{molar mass of M(OH)}_2] - [\text{molar mass of (OH)}_2]$$

ANSWER:

63.54 $\mathrm{g/mol}$

ANSWER:

Copper

Correct

The hydroxide formula is $\mathrm{Cu(OH)_2}$.

Limiting Reactant Procedure

In the following chemical reaction, 2 mol of A will react with 1 mol of B to produce 1 mol of $\mathrm{A_2B}$ without anything left over:

$$2\mathrm{A} + \mathrm{B} \rightarrow \mathrm{A_2B}$$

But what if you're given 2.8 mol of A and 3.2 mol of B? The amount of product formed is limited by the reactant that runs out first, called the limiting reactant. To identify the limiting reactant, calculate the amount of product formed from each amount of reactant separately:

$$2.8 \ \cancel{\text{mol A}} \times \frac{1 \text{ mol A}_2\text{B}}{2 \ \cancel{\text{mol A}}} = 1.4 \text{ mol A}_2\text{B}$$

$$3.2 \ \cancel{\text{mol B}} \times \frac{1 \text{ mol A}_2\text{B}}{1 \ \cancel{\text{mol B}}} = 3.2 \text{ mol A}_2\text{B}$$

Notice that less product is formed with the given amount of reactant A. Thus, A is the limiting reactant, and a maximum of 1.4 mol of $\mathrm{A_2B}$ can be formed from the given amounts.

Aluminum reacts with chlorine gas to form aluminum chloride via the following reaction:

$$2\mathrm{Al}(s) + 3\mathrm{Cl_2}(g) \rightarrow 2\mathrm{AlCl_3}(s)$$

You are given 16.0 g of aluminum and 21.0 g of chlorine gas.

Part A

If you had excess chlorine, how many moles of of aluminum chloride could be produced from 16.0 g of aluminum?

Express your answer to three significant figures and include the appropriate units.

Hint 1. Determine the molar mass of aluminum

Use the periodic table to determine the molar mass of aluminum.

Express your answer to four significant figures and include the appropriate units.

ANSWER:

$26.98\ \frac{\mathrm{g}}{\mathrm{mol}}$

Correct

Hint 2. Convert mass to moles

Convert 16.0 g of aluminum into moles.

Express your answer to three significant figures and include the appropriate units.

Hint 1. Converting mass to moles

Use the molar mass of aluminum as a conversion factor to calculate the number of moles. Recall that molar mass is in units of $\mathrm{g/mol}$, and your answer should be in units of mol.

ANSWER:

$0.593\ \mathrm{mol}$

Correct

Now use the mole ratio of Al to $\mathrm{AlCl_3}$ to determine the amount of product that will form.

ANSWER:

$0.593\ \mathrm{mol}$

Correct

Part B

If you had excess aluminum, how many moles of aluminum chloride could be produced from 21.0 g of chlorine gas, $\mathrm{Cl_2}$?

Express your answer to three significant figures and include the appropriate units.

Hint 1. Determine the molar mass of chlorine

Use the periodic table to calculate the molar mass of chlorine, $\mathrm{Cl_2}$.

Express your answer to four significant figures and include the appropriate units.

ANSWER:

$70.91\ \frac{\mathrm{g}}{\mathrm{mol}}$

Correct

Hint 2. Convert mass to moles

Convert 21.0 g of $\mathrm{Cl_2}$ to moles.

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answer to three significant figures and include the appropriate units.

Hint 1. Converting mass to moles

Use the molar mass of chlorine as a conversion factor to calculate moles. Recall that molar mass is in units of $\mathrm{g/mol}$, and your answer should be in units of mol.

ANSWER:

0.296 mol

Correct

Now use the mole ratio of $\mathrm{Cl_2}$ to $\mathrm{AlCl_3}$ to determine the amount of product that will form.

ANSWER:

0.197 mol

Correct

By comparing your answers for Parts A and B, you can determine which reactant is limiting. Keep in mind that the limiting reactant is the one that produces the lesser amount of product.

Part C

Aluminum reacts with chlorine gas to form aluminum chloride via the following reaction:

$$2\mathrm{Al}(s) + 3\mathrm{Cl_2}(g) \rightarrow 2\mathrm{AlCl_3}(s)$$

What is the maximum mass of aluminum chloride that can be formed when reacting 16.0 g of aluminum with 21.0 g of chlorine?

Express your answer to three significant figures and include the appropriate units.

Hint 1. Determine the maximum number of moles

A mass of 16.0 g of aluminum with excess chlorine produces 0.593 mol of product, whereas 21.0 g of chlorine with excess aluminum produces only 0.197 mol of product. How much product will form when 16.0 g of aluminum reacts with 21.0 g of chlorine?

ANSWER:

- 0.593 mol of product
- 0.197 mol of product

Hint 2. Determine the molar mass of the product

What is the molar mass of aluminum chloride, $\mathrm{AlCl_3}$?

Express your answer to four significant figures and include the appropriate units.

ANSWER:

133.3 $\frac{\mathrm{g}}{\mathrm{mol}}$

ANSWER:

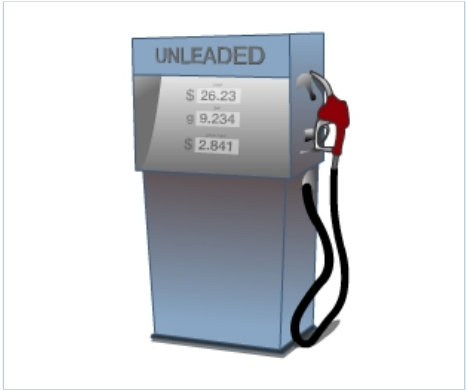
26.3 g

Correct

± Limiting Reactants

The octane rating of gasoline is a relationship of the burning efficiency of the given gasoline mixture to the burning efficiency of octane ($\mathrm{C_8H_{18}}$). Like most hydrocarbons, octane reacts with oxygen gas to produce carbon dioxide and water. The unbalanced equation for this reaction is

$$\mathrm{C_8H_{18}}(g) + \mathrm{O_2}(g) \rightarrow \mathrm{CO_2}(g) + \mathrm{H_2O}(g)$$



Part A

Determine the balanced chemical equation for this reaction.

$C_8H_{18}\,(g) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(g)$

Enter the coefficients for each compound in order, separated by commas. For example, `1,2,3,4` would indicate one mole of C_8H_{18}, two moles of O_2, three moles of CO_2, and four moles of H_2O.

Hint 1. Balance the carbon atoms

How many units of CO_2 are necessary to balance the carbon atoms in one unit of C_8H_{18}?

Enter the number of units of CO_2 as an integer.

ANSWER:

CO_2 8 units

Hint 2. Balance the hydrogen atoms

How many units of H_2O are necessary to balance the carbon atoms in one unit of C_8H_{18}?

Enter the number of units of H_2O as an integer.

ANSWER:

H_2O 9 units

Hint 3. Balance the oxygen atoms

How many units of O_2 are necessary to balance the total number of oxygen atoms in $8CO_2$ and $9H_2O$? Hint: It might not be a whole number.

Enter the number of units of O_2 as a fraction or decimal.

ANSWER:

25/2

Hint 4. Multiply to get whole numbers

At this point, our equation looks like this:

$C_8H_{18}\,(g) + 12.5\,O_2\,(g) \rightarrow 8\,CO_2\,(g) + 9\,H_2O\,(g)$

It is currently balanced, but the coefficients should be integers, not fractions or decimals. Therefore, you need to double each coefficent.

ANSWER:

2,25,16,18

Correct

It is important to balance a chemical equation before using it for calculations. Checking that equations are balanced will help you avoid many errors in chemistry problems.

Balanced chemical equation

$2\,C_8H_{18}\,(g) + 25\,O_2\,(g) \rightarrow 16\,CO_2\,(g) + 18\,H_2O\,(g)$

Part B

0.320 mol of octane is allowed to react with 0.810 mol of oxygen. Which is the limiting reactant?

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Hint 1. How to approach the problem

Method 1: Calculate the number of moles of product that would be formed from the given number of moles of each reactant. The reactant that yields less product is the limiting reactant.

Method 2: Divide the number of moles of each reactant by its coefficient in the balanced equation. The reactant with the smallest moles/coefficient ratio is the limiting one.

Hint 2. Method 1

Which reactant produces the least amount of CO_2?

Hint 1. Find the number of moles produced by octane

How many moles of CO_2 could be produced from 0.320 mol of octane?

Express your answer with the appropriate units.

ANSWER:

2.56 mol

Hint 2. Find the number of moles produced by oxygen

How many moles of CO_2 could be produced from 0.810 mol of oxygen?

Express your answer with the appropriate units.

ANSWER:

0.518 mol

ANSWER:

- octane
- oxygen

Hint 3. Method 2

Which reactant has the smallest moles/coefficient ratio?

Hint 1. Find the ratio for octane

For octane: $\frac{\text{moles}}{\text{coefficient}} = ?$

Enter the moles/coefficient ratio for octane

ANSWER:

? = 0.160

Hint 2. Find the ratio for oxygen

For oxygen: $\frac{\text{moles}}{\text{coefficient}} = ?$

Enter the moles/coefficient ratio for oxygen

ANSWER:

? = 3.24×10^{-2}

ANSWER:

- octane
- oxygen

ANSWER:

Typesetting math: 68%

- octane
- oxygen

Correct

Now that you have identified oxygen as the limiting reactant, you can use the number of moles of oxygen to find the numbers of moles of all the other substances.

Part C

How many moles of water are produced in this reaction?

Express your answer with the appropriate units.

Hint 1. How to approach the problem

Take the number of moles of oxygen and divide it by the coefficient for oxygen. Then multiply by the coefficient for water.

ANSWER:

$\mathrm{H_2O}$ produced = 0.583 mol

Correct

Part D

After the reaction, how much octane is left?

Express your answer with the appropriate units.

Hint 1. How to approach the problem

Use the number of moles of oxygen to calculate the number of moles of octane that were acutally used. Then subtract that amount from the given number of moles of octane.

Hint 2. Find the number of moles of reacted octane

How much octane can react with 0.810 mol of oxygen?

Express your answer with the appropriate units.

ANSWER:

number of moles of $\mathrm{C_8H_{18}}$ reacted = 6.48×10^{-2} mol

ANSWER:

moles of $\mathrm{C_8H_{18}}$ remaining = 0.255 mol

Correct

When hydrocarbons burn with too little oxygen present, they produce carbon monoxide gas in addition to the water and carbon dioxide shown in this reaction. Carbon monoxide is much more toxic than carbon dioxide. The production of carbon monoxide by improperly ventilated automobile engines and other hydrocarbon-burning apparatus has resulted in numerous deaths.

± Percent Yield

The Haber-Bosch process is a very important industrial process. In the Haber-Bosch process, hydrogen gas reacts with nitrogen gas to produce ammonia according to the equation

$$3\mathrm{H_2(g)} + \mathrm{N_2(g)} \rightarrow 2\mathrm{NH_3(g)}$$

The ammonia produced in the Haber-Bosch process has a wide range of uses, from fertilizer to pharmaceuticals. However, the production of ammonia is difficult, resulting in lower yields than those predicted from the chemical equation.

1.48 $\mathrm{g\ H_2}$ is allowed to react with 10.4 $\mathrm{g\ N_2}$, producing 1.48 $\mathrm{g\ NH_3}$.

Part A

What is the theoretical yield for this reaction under the given conditions?

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Express your answer to three significant figures and include the appropriate units.

Hint 1. How to approach the problem

First, find the limiting reactant. Then, use stoichiometry to convert the mass of that reactant to the mass of ammonia, the theoretical yield.

Hint 2. Find the limiting reactant

What is the limiting reactant based on the given masses?

Hint 1. Find the number of moles of hydrogen

How many moles of $\mathrm{H_2}$ are present?

Express your answer to three significant figures and include the appropriate units.

Hint 1. Determine the molar mass of hydrogen

What is the molar mass of hydrogen, $\mathrm{H_2}$? Use the periodic table as needed.

Express your answer to four significant figures and include the appropriate units.

ANSWER:

$2.016\ \frac{\mathrm{g}}{\mathrm{mol}}$

ANSWER:

$0.734\ \mathrm{mol}$

Hint 2. Find the number of moles of nitrogen

How many moles of $\mathrm{N_2}$ are present?

Express your answer to three significant figures and include the appropriate units.

Hint 1. Determine the molar mass of nitrogen

What is the molar mass of hydrogen, $\mathrm{N_2}$? Use the periodic table as needed.

Express your answer to four significant figures and include the appropriate units.

ANSWER:

$28.02\ \frac{\mathrm{g}}{\mathrm{mol}}$

ANSWER:

$0.372\ \mathrm{mol}$

ANSWER:

- (•) hydrogen
- ( ) nitrogen
- ( ) ammonia

Hint 3. Identify the factor to convert from the limiting reactant to the product

Based on the limiting reactant, which of the following would be the correct conversion factor to use to convert moles of the limiting reactant to moles of product that could form?

ANSWER:

- $\frac{2\ \text{mol}\ NH_3}{3\ \text{mol}\ H_2}$
- $\frac{1\ \text{mol}\ NH_3}{1\ \text{mol}\ H_2}$
- $\frac{3\ \text{mol}\ H_2}{2\ \text{mol}\ VH_3}$
- $\frac{1\ \text{mol}\ N_2}{2\ \text{mol}\ NH_3}$
- $\frac{2\ \text{mol}\ NH_3}{1\ \text{mol}\ N_2}$
- $\frac{1\ \text{mol}\ NH_3}{1\ \text{mol}\ N_2}$

ANSWER:

8.34 g

Correct

The theoretical yield is the amount you would get if the reaction went to completion giving a yield of 100%. In practice, 100% yield is difficult to achieve because the experimental techniques used to isolated the product are not completely efficient.

Part B

What is the percent yield for this reaction under the given conditions?

Express your answer to three significant figures and include the appropriate units.

Hint 1. The definition of percent yield

Recall that

$$\text{percent yield} = \frac{\text{actual yield (g)}}{\text{theoretical yield (g)}} \times 100\%$$

ANSWER:

17.8 %

Correct

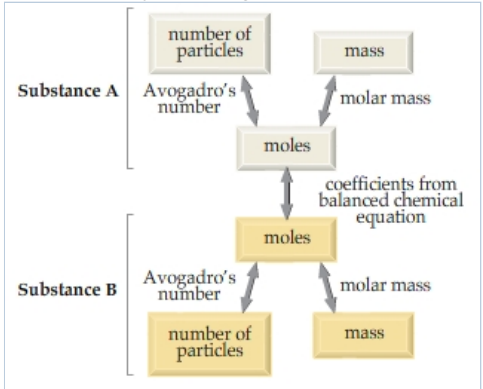
Industrially, this reaction is carried out using a catalyst, and the ammonia is constantly removed as it is produced to increase the yield of the reaction. Economically, the percent yield is a very important consideration.

± Learning Stoichiometry

Learning Goal:

To understand how to use stoichiometry to convert between quantities of reactants and products in chemical equations.

Stoichiometry describes the quantitative relationships among the reactants and products of a balanced reaction by directly comparing mole ratios. Stoichiometry can be used to convert mass, number of moles, or number of particles between products and reactants, as shown in the flowchart displayed in the figure.



Converting between quantities

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To convert from a given quantity of one reactant or product to the quantity of another reactant or product:

- First, convert the given quantity to moles. Use molar masses to convert masses to moles, and use Avogadro's number (6.02×10^{23} particles per mole) to convert number of particles to moles.
- Next, convert moles of the given reactant or product to moles of the desired reactant or product using the coefficients of the balanced chemical equation. For example, in the chemical equation
$$2H_2 + O_2 \rightarrow 2H_2O$$
the coefficients tell us that 2 mol of H_2 reacts with 1 mol of O_2 to produce 2 mol of H_2O.
- Finally, convert moles of the desired reactant or product back to the desired units. Again, use molar masses to convert from moles to masses, and use Avogadro's number to convert from moles to number of particles.

Reaction of hydrogen and nitrogen to form ammonia

Hydrogen gas, H_2, reacts with nitrogen gas, N_2, to form ammonia gas, NH_3, according to the equation

$$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$$

NOTE: Throughout this tutorial use molar masses expressed to five significant figures.

Part A

How many moles of NH_3 can be produced from 19.5 mol of H_2 and excess N_2?

Express your answer numerically in moles.

Hint 1. How to approach the problem

First determine the ratio between the number of moles of H_2 consumed and the number of moles of NH_3 produced. Then, apply this ratio to the number of moles of H_2 to find the amount of product formed.

Hint 2. Apply the molar ratio

Which of the following correctly uses the coefficients in the balanced chemical equation to find the number of moles of NH_3 produced given the number of moles of H_2.

ANSWER:

- Multiply by $\frac{3\ \mathrm{mol\ H_2}}{2\ \mathrm{mol\ NH_3}}$.
- Multiply by $\frac{2\ \mathrm{mol\ H_2}}{3\ \mathrm{mol\ NH_3}}$.
- Multiply by $\frac{3\ \mathrm{mol\ NH_3}}{2\ \mathrm{mol\ H_2}}$.
- Multiply by $\frac{2\ \mathrm{mol\ NH_3}}{3\ \mathrm{mol\ H_2}}$.

ANSWER:

13.0 $\mathrm{mol\ NH_3}$

Correct

Part B

How many grams of NH_3 can be produced from 3.82 mol of N_2 and excess H_2.

Express your answer numerically in grams.

Hint 1. How to approach the problem

The amount of nitrogen gas available in the reaction is given in moles. Using the coefficients of the chemical equation, determine how many moles of NH_3 can be produced. Next, use the molar mass of NH_3 to find the corresponding mass in grams of ammonia.

Hint 2. Calculate the number of moles of ammonia produced

How many moles of ammonia can be produced from 3.82 mol of N_2 based on the balanced reaction?

Express your answer numerically in moles.

ANSWER:

7.64 $\mathrm{mol\ NH_3}$

Hint 3. Determine the molar mass of ammonia

What is the molar mass of ammonia, NH_3?

Express your answer in grams per mole using five significant figures.

ANSWER:

17.031 $\mathrm{g/mol}$

ANSWER:

130 $\mathrm{g\ NH_3}$

Correct

Part C

How many grams of $\mathrm{H_2}$ are needed to produce 10.28 g of $\mathrm{NH_3}$?

Express your answer numerically in grams.

Hint 1. How to approach the problem

The mass of ammonia produced by the reaction is given in grams. First, convert this mass to moles. Use the computed number of moles of ammonia to find the required number of moles of hydrogen using the balanced reaction. Finally, convert the number of moles of hydrogen to the mass of hydrogen in grams using the molar mass of hydrogen.

Hint 2. Calculate the number of moles of ammonia

How many moles of $\mathrm{NH_3}$ are in 10.28 g of $\mathrm{NH_3}$?

Express your answer numerically in moles.

Hint 1. How to convert mass of ammonia to moles

To convert the mass of ammonia to moles, divide the mass in grams by the molar mass in grams per mole.

ANSWER:

0.6036 $\mathrm{mol\ NH_3}$

Hint 3. Calculate the number of moles of hydrogen

How many moles of $\mathrm{H_2}$ are needed to produce 10.28 g of $\mathrm{NH_3}$?

Express your answer numerically in moles.

ANSWER:

0.9054 $\mathrm{mol\ H_2}$

Hint 4. Determine the molar mass of hydrogen gas

What is the molar mass of $\mathrm{H_2}$?

Express your answer numerically in grams per mole to five significant figures.

ANSWER:

2.0159 $\mathrm{g/mol}$

ANSWER:

1.825 $\mathrm{g\ H_2}$

Correct

Part D

How many molecules (not moles) of $\mathrm{NH_3}$ are produced from 3.32×10^{-4} g of $\mathrm{H_2}$?

Express your answer numerically as the number of molecules.

Hint 1. How to approach the problem

The mass of available hydrogen gas is given in grams. First, convert this mass to moles using the molar mass of hydrogen gas. Use the computed number of moles of hydrogen gas to find the number of moles of ammonia produced in the reaction. Finally, convert the number of moles of ammonia to the number of molecules of ammonia using Avogadro's number.

Hint 2. Calculate the number of moles of hydrogen gas

How many moles of $\mathrm{H_2}$ are in 3.32×10^{-4} g of $\mathrm{H_2}$?

Express your answer numerically in moles.

ANSWER:

1.65×10^{-4} mol $\mathrm{H_2}$

Hint 3. Calculate the number of moles of ammonia produced

How many moles of $\mathrm{NH_3}$ can be produced from 3.32×10^{-4} g of $\mathrm{H_2}$?

Express your answer numerically in moles.

ANSWER:

1.10×10^{-4} mol $\mathrm{NH_3}$

ANSWER:

6.61×10^{19} molecules

Correct

± Concentration

Molarity (M) is defined as the number of moles of solute divided by the solution volume expressed in liters:

$$\text{molarity} = \frac{\text{moles of solute}}{\text{volume of solution (L)}}$$

For example, 1 M HCl contains 1 mol of HCl dissolved in 1 L of the water. When a concentrated solution is diluted, the number of moles of solute stays constant; only the volume of the solution is changed. A dilution indicates an increase in solution volume and, therefore, the concentration of the solution must decrease. If you add more water to the HCl solution considered above, so that now the volume is 2 L , the number of moles remains the same but the volume is doubled. Hence the molarity of the solution is now 1 mol in a 2 L solution, that is, $(1/2)$ M or 0.5 M.

The number of moles of solute before and after dilution can be calculated by multiplying molarity times volume. We can set up the following equations:

$$\begin{aligned}\text{moles of solute} &= \text{molarity} \times \text{volume} \\ &= M_\mathrm{i} \times V_\mathrm{i} = M_\mathrm{f} \times V_\mathrm{f}\end{aligned}$$

where M_i is the initial molarity (of the concentrated solution), V_i is the initial volume, M_f is the final molarity (of the diluted solution), and V_f is the final volume.

In the HCl solution example the initial molarity is 1 M, the initial volume is 1 L, the final volume is 2 L, and the molarity is 0.5 M. Thus the number of moles present in these solutions is

$$M_\mathrm{i}V_\mathrm{i} = M_\mathrm{f}V_\mathrm{f} = 1\ M \times 1\ \mathrm{L} = 0.5\ M \times 2\ \mathrm{L} = 1\ \mathrm{mol}$$

Part A

The following five beakers, each containing a solution of sodium chloride (NaCl, also known as table salt), were found on a lab shelf:

Beaker	Contents
1	200. mL of 1.50 M NaCl solution
2	100. mL of 3.00 M NaCl solution
3	150. mL of solution containing 25.5 g of NaCl
4	100. mL of solution containing 25.5 g of NaCl
5	300. mL of solution containing 0.450 mol NaCl

Arrange the solutions in order of decreasing concentration.

Rank from most concentrated to least concentrated. To rank items as equivalent, overlap them.

Hint 1. How to approach the problem

To arrange the beakers from highest concentration to lowest concentration, you need to know the molar concentration of each solution. Molar concentration of the solution is the concentration of solution expressed in terms of molarity M.

The molar concentrations of beakers 1 and 2 are given in the table. You need to calculate the molar concentrations of beakers 3, 4, and 5 using the equation

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

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For each beaker, make sure to convert the amount of sodium chloride to moles and the volume of solution to liters before calculating the molarity of the solution.

Hint 2. Calculate the molarity of the solution in beaker 3

A 150.-mL sample of the solution contains 25.5 g of NaCl. Calculate the molarity of the solution.

Express your answer to three significant figures and include the appropriate units.

Hint 1. Calculate the number of moles of solute, NaCl

How many moles are present in 25.5 g of NaCl? Use the molar mass to find the value.

Express your answer to three significant figures and include the appropriate units.

Hint 1. Calculate the molar mass of sodium chloride

Molar mass is the mass of the substance that is present in 1 mol of it. For example, the molar mass of HCl is the atomic mass of H plus the atomic mass of Cl, which equals $1 + 35.5 = 36.5\ \mathrm{g/mol}$.

What is the molar mass of NaCl? You can use the periodic table to calculate the molar mass of NaCl.

Express your answer to four significant figures and include the appropriate units.

ANSWER:

$58.44\ \frac{\mathrm{g}}{\mathrm{mol}}$

Hint 2. Determine how to convert from grams to moles

When calculating the number of moles from grams, which of the following conversion factors will allow the units "grams" (\rm g) to cancel?

ANSWER:

- \large{\rm Given~mass~of~\rm NaCl~in~g\times \frac{1~\rm mol~NaCl}{molar ~mass~ of~\rm NaCl~in~g}}
- \large{\rm Given~mass~of~\rm NaCl~in~g\times \frac {molar~mass~of~ \rm NaCl~in~g}{1~\rm mol~NaCl}}

ANSWER:

0.436 {\rm mol}

Hint 2. Calculate the volume of the solution in liters

In the calculation for the molarity the volume of the solution is in liters. Convert the given volume of the solution, 150. \rm mL, to liters. Consider that there are 1000 \rm mL in 1 \rm L.

Express your answer to three significant figures and include the appropriate units.

ANSWER:

0.150 {\rm L}

ANSWER:

\rm molarity_3 = 2.91 {\rm M}

Hint 3. Calculate the molarity of the solution in beaker 4

A 100.-\rm mL sample of the solution contains 25.5 {\rm g} of \rm NaCl. Calculate the molarity of the solution.

Express your answer to three significant figures and include the appropriate units.

Hint 1. Calculate the number of moles of solute

How many moles are in 25.5 {\rm g} of \rm NaCl?

Express your answer to three significant figures and include the appropriate units.

Hint 1. Calculate the molar mass of sodium chloride

Molar mass is the mass of the substance that is present in 1 \rm mol of it. For example, the molar mass of \rm HCl is the atomic mass of \rm H plus the atomic mass of \rm Cl, which equals 1 + 35.5 = 36.5 ~\rm g/mol.

the molar mass of \rm NaCl? You can use the periodic table to calculate the molar mass of \rm NaCl.

Express your answer to four significant figures and include the appropriate units.

ANSWER:

58.44 \large{{\rm \frac{g}{mol}}}

Hint 2. Determine how to convert from grams to moles

When calculating the number of moles from grams, which of the following conversion factors will allow the units "grams" (\rm g) to cancel?

ANSWER:

- \large{\rm Given~mass~ of~\rm NaCl~in~g\times \frac {1~ \rm mol~NaCl}{molar~mass~of~\rm NaCl~in~g}}
- \large{\rm Given~mass~of~\rm NaCl~in~g\times \frac {molar~mass~of~\rm NaCl~in~g}{1~ \rm mol~NaCl}}

ANSWER:

0.436 {\rm mol}

Hint 2. Calculate the volume of the solution in liters

In the calculation for the molarity the volume of the solution is in liters. Convert the given volume of the solution, 150. \rm mL, to liters. Consider that there are 1000 \rm mL in 1 \rm L.

Express your answer to three significant figures and include the appropriate units.

ANSWER:

0.100 {\rm L}

ANSWER:

\rm molarity_4 = 4.36 {\rm M}

Hint 4. Calculate the molarity of the solution in beaker 5

A 300.-\rm mL volume of a \rm NaCl solution contains 0.450 \rm mol of solute. What is the molarity of this solution?

Express your answer to three significant figures and include the appropriate units.

Hint 1. Convert the volume of the solution to liters

In the calculation for the molarity the volume of the solution is in liters. Convert the given volume of the solution, 300. \rm mL, to liters. Consider that there are 1000 \rm mL in 1 \rm L.

Express your answer to three significant figures and include the appropriate units.

ANSWER:

0.300 {\rm L}

ANSWER:

\rm molarity_5 = 1.50 {\rm M}

ANSWER:

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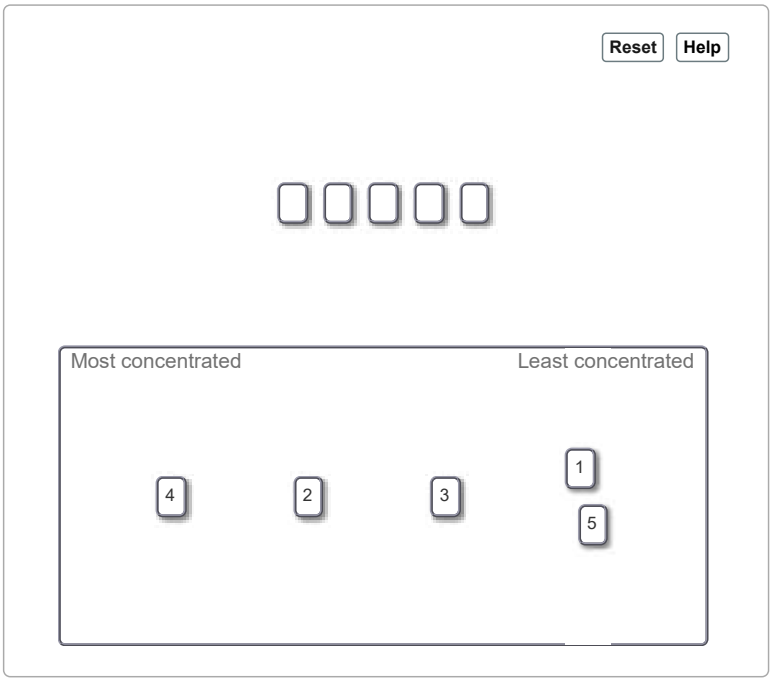

Correct

Molarity (or molar concentration) is one of several possible ways to express the solution concentration. Since molarity refers to the number of moles of solute, it is very useful in stoichiometry calculations for reactions taking place in solution. The unit of molarity, M, can also be written as \rm mol/L.

Part B

A student placed 18.0 {\rm g} of glucose (\rm C_6H_{12}O_6) in a volumetric flask, added enough water to dissolve the glucose by swirling, then carefully added additional water until the 100. \rm mL mark on the neck of the flask was reached. The flask was then shaken until the solution was uniform. A 60.0 \rm mL sample of this glucose solution was diluted to 0.500 \rm L. How many grams of glucose are in 100. \rm mL of the final solution?

Express your answer to three significant figures and include the appropriate units.

Hint 1. How to approach the problem

Start by calculating the molar concentration of the solution when 18.0 {\rm g} glucose is dissolved in 100. \rm mL water. Consider this solution as your initial solution. Consider the molarity of the initial solution as \rm X M. Then use the dilution equation to find the concentration of the final solution. The final solution forms when you take 45 \rm mL of \rm X M solution and dilute it to 0.5 \rm L. Here using the dilution equation you will find the molarity of the final solution. The final solution is the one that has a volume of 0.5 \rm L.

Once the molar concentration of the final solution is known, use it to find the number of moles of solute in 100. \rm mL of the final solution. Finally, convert moles to grams using the molar mass of glucose.

Hint 2. Calculate the molar mass of glucose

Calculate the molar mass of glucose, \rm C_6H_{12}O_6.

Express your answer to five significant figures and include the appropriate units.

Hint 1. How to calculate the molar mass

1. Look up the atomic masses of carbon, hydrogen, and oxygen in the periodic table.
2. Determine the sum of the masses of 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms.
3. Note that the molecular mass in units of \rm amu is equivalent to the molar mass in \rm g/mol.

ANSWER:

180.16 \large{{\rm \frac{g}{mol}}}

Hint 3. Calculate the molar concentration of the initial solution

Calculate the molar concentration of the initial glucose (\rm C_6H_{12}O_6) solution. Recall that 100. \rm mL of the solution contains 18.0 {\rm g} of glucose.

Express your answer to four significant figures and include the appropriate units.

Hint 1. Calculate the number of moles of glucose

Convert 18.0 {\rm g} of glucose (\rm C_6H_{12}O_6) to moles.

Express your answer to four significant figures and include the appropriate units.

Hint 1. Determine the conversion factor to convert grams to moles

When calculating the number of moles, given the mass in grams, which of the following conversion factors will allow the units "grams" (\rm g) to cancel?

ANSWER:

- \large{\rm Given~mass~of~\rm C_6H_{12}O_6~in~g\times \frac{1~\rm mol~C_6H_{12}O_6}{molar~mass~of~\rm C_6H_{12}O_6~in~g}}
- \large{\rm Given~mass~of~\rm C_6H_{12}O_6~in~g\times \frac{molar~mass~of~\rm C_6H_{12}O_6~in~g}{1~\rm mol~C_6H_{12}O_6}}

ANSWER:

9.991×10^{-2} {\rm mol}

ANSWER:

\texttip{M_{\rm i}}{M_i} = 0.9991 {\rm M}

Hint 4. Calculate the concentration of the final solution

A 60.0 \rm mL sample of a 0.9991 M solution of glucose was diluted to 0.500 \rm L. Calculate the concentration of the final solution. Use the dilution equation to find the molarity of the final solution.

Express your answer to four significant figures and include the appropriate units.

Hint 1. Rearrange the dilution equation to solve for \texttip{M_{\rm f \hspace{1 pt}}}{M_f}

Rearrange the dilution equation

$$M_{\rm i}\times V_{\rm i}=M_{\rm f \hspace{1 pt}}\times V_{\rm f \hspace{1 pt}}$$

to solve for \texttip{M_{\rm f \hspace{1 pt}}}{M_f}.

ANSWER:

- M_{\rm i}\times V_{\rm i}=M_{\rm f \hspace{1 pt}}\times V_{\rm f \hspace{1 pt}}
- M_{\rm f \hspace{1 pt}}=M_{\rm i}\times V_{\rm i}\times V_{\rm_f}
- \large{M_{\rm f \hspace{1 pt}}=\frac{M_{\rm i}\times V_{\rm i}}{V_{\rm f \hspace{1 pt}}}}

ANSWER:

\texttip{M_{\rm f \hspace{1 pt}}}{M_f} = 0.1199 {\rm M}

Hint 5. Calculate the number of moles in 100. \rm mL of the final solution

How many moles of glucose are present in 100. \rm mL of a 0.1199 M glucose solution?

Express your answer to four significant figures and include the appropriate units.

Hint 1. Rearrange the molarity equation to determine the number of moles

Recall that molarity is equal to moles divided by volume in liters:

$$\large{\rm molarity=\frac{\rm moles~of~solute}{\rm volume~of~solution~(L)}}$$

Solve this expression for moles.

Express your answer in terms of molarity and volume.

ANSWER:

M volume

ANSWER:

1.199×10^{-2} {\rm mol}

Typesetting math: 68%

ANSWER:

2.16 {\rm g}

Correct

Another way to approach this problem is to consider that 60.0 {\rm mL} is 60.0 \% of the initial 100. \rm mL volume and so 60.0 \% of the original mass of glucose was present in the sample that got diluted to 0.500 \rm L. Then, we know that one-fifth of 0.5 \rm L solution (100. \rm mL) was taken. One-fifth of 60.0 \% of 18.0 {\rm g} is equal to the final answer.

± Molarity

The chemical 5-amino-2,3-dihydro-1,4-phthalazinedione, better known as luminol, is used by forensic scientists in analyzing crime scenes for the presence of washed-away blood. Luminol is so sensitive that it can detect blood that has been diluted 10,000 times. A basic solution of luminol is often sprayed onto surfaces that are suspected of containing minute amounts of blood.

Luminol has a molecular weight of 177 \rm g/mol.

Part A

The forensic technician at a crime scene has just prepared a luminol stock solution by adding 11.0 {\rm g} of luminol into a total volume of 75.0 \rm mL of \rm H_2O.

What is the molarity of the stock solution of luminol?

Express your answer with the appropriate units.

Hint 1. How to approach the problem

In solving this problem, first calculate how many moles of luminol have been used to make the solution, using the molecular weight of luminol to convert grams to moles of luminol. Then, divide the number of moles by the volume of the solution, keeping in mind that molarity is reported as moles per liter.

Hint 2. Definition of molarity

Molarity, M, is the number of moles of solute per liter of solution: M = {\rm moles/ liter}. The units for molarity, M, are moles per liter.

Hint 3. Determine the number of moles of luminol in the stock solution

How many moles of luminol are present in the stock solution?

Express your answer with the appropriate units.

Hint 1. How to determine moles from grams

To convert from grams to moles of a compound, divide the number of grams by the molecular weight of the compound in grams per mole. The molecular weight of luminol is given in the problem description.

ANSWER:

moles of luminol = 6.21×10^{-2} {\rm mol}

ANSWER:

molarity of luminol solution = 0.829 {\rm M}

Correct

Part B

Before investigating the scene, the technician must dilute the luminol solution to a concentration of 6.00×10^{-2} {\rm \it M} . The diluted solution is then placed in a spray bottle for application on the desired surfaces.

How many moles of luminol are present in 2.00 L of the diluted spray?

Express your answer with the appropriate units.

Hint 1. How to approach the problem

M = {moles/ liter}. The formula is the same, but this time you are solving for moles.

ANSWER:

moles of luminol = 0.120 {\rm mol}

Correct

Part C

What volume of the stock solution (Part A) would contain the number of moles present in the diluted solution (Part B)?

Express your answer with the appropriate units.

Hint 1. How to approach the problem

M = {\rm moles/ liter}. The formula is the same, but this time you are solving for volume. Use the molarity in part A, 0.829 {\rm M} , and the number of moles in part B, 0.120 {\rm \rm mol} , for the other variables. Substitute these values into the equation, and rearrange to solve for the volume.

ANSWER:

volume = 145 {\rm mL}

Correct

To make the diluted solution, the technician would use this volume of the concentrated stock solution and add water until a total volume of 2.00 \rm L was reached.

Give It Some Thought 14.3

Part A

How do reaction rate, rate law, and rate constant differ?

Drag the terms on the left to the appropriate blanks on the right to complete the sentences.

ANSWER:

Reset Help

reaction rate
rate law
rate constant

1. Reaction rate is the change in the concentration of a reactant (disappearance) or product (appearance) over a time interval.
2. Rate constant is part of the rate law .
3. Rate law governs how the reaction rate depends on the concentrations of the reactants.

Correct

Magnitude of the Equilibrium Constant

For a reaction

$$a{\rm A} + b{\rm B} \rightleftharpoons c{\rm C} + d{\rm D}$$

the equilibrium constant \texttip{K}{K} is defined as the ratio of products to reactants at equilibrium:

$$\large{K = \frac{[{\rm C}]^c[{\rm D}]^d}{[{\rm A}]^a[{\rm B}]^b}}$$

Because \texttip{K}{K} represents the ratio of products to reactants, the magnitude of \texttip{K}{K} is an indicator of the levels of products and reactants present when the reaction is at equilibrium. Some chemical reactions proceed almost fully to product, whereas other chemicals hardly react with each other at all, or react so as to have significant amounts of both reactants and products at equilibrium.

Part A

Hydrogen is used as a rocket fuel because it is very light and reacts explosively and completely with oxygen. For the combustion reaction

$${\rm 2\ H}_{2}\ ({\rm g}) + {\rm O}_{2}\ ({\rm g}) \rightleftharpoons 2\ {\rm H_2O}\ ({\rm g})$$

what is the likely magnitude of the equilibrium constant \texttip{K}{K}?

Hint 1. How to approach the problem

The equilibrium constant is the ratio of product to reactants at equilibrium. Think about the relative amounts of products versus reactants in a reaction that goes essentially to completion and how this affects the size of \texttip{K}{K}.

Typesetting math: 68%

ANSWER:

- K < 10^{-3}
- 10^{-3} < K < 10^{3}
- K =0
- K > 10^{3}

Correct

The actual \texttip{K}{K} for the combustion reaction of hydrogen gas with oxygen at room temperature is K = 1.4 \times 10^{83}!

Part B

The industrial production of lime (\rm CaO) from calcium carbonate is accomplished via the following reaction:

$$\rm CaCO_3\,(s) \rightleftharpoons CaO\,(s) + CO_2\,(g)$$

Given the following data:

Temperature (\rm K)	\texttip{K}{K}
298	1.93\times 10^{-23}
1200	1.01

what can be said about this reaction?

Hint 1. How to approach the problem

Recall that the equilibrium constant is the ratio of product to reactants at equilibrium. Think about the relative amounts of products versus reactants at different values of \texttip{K}{K}.

Hint 2. Interpret the meaning of a small *K*

Recall that the value of \texttip{K}{K} is small for reactions that form very little product. Reactions that form very little product are said to _____.

ANSWER:

- have an equilibrium that lies far to the left
- have \texttip{K}{K} values in excess of 1000
- have an equilibrium that lies far to the right
- result in very little leftover reactants at equilibrium

ANSWER:

- Lower temperatures result in more lime formation.
- The reaction makes more lime at higher temperatures.
- The reaction goes to completion at 1200 \rm K.
- The equilibrium lies far to the right at room temperature.

Correct

Part C

For the reaction

$$\rm COCl_2(g) \rightleftharpoons CO(g) + Cl_2(g),\ {\it K} = \rm 2.19\times10^{-10}\ at\ 373\ \rm K$$

what can be said about this reaction at this temperature?

Hint 1. How to approach the problem

Recall that the equilibrium constant is the ratio of product to reactants at equilibrium. Think about the relative amounts of products versus reactants at small values of \texttip{K}{K}.

Hint 2. Interpreting the *K* value

Typesetting math: 68% 10^{-3} is considered small, meaning that the reaction does not contain a significant amount of product at equilibrium.

ANSWER:

- The equilibrium lies far to the right.
- The reaction will proceed very slowly.
- The reaction contains significant amounts of products and reactants at equilibrium.
- (•) The equilibrium lies far to the left.

Correct

Part D

For the reaction

$$\rm HCONH_2(g) \rightleftharpoons NH_3(g) + CO(g), {\it K} = \rm 4.84 \text{ at } 400 \rm K$$

what can be said about this reaction at this temperature?

Hint 1. How to approach the problem

Recall that the equilibrium constant is the ratio of product to reactants at equilibrium. Think about the relative amounts of products versus reactants at small values of \texttip{K}{K}.

Hint 2. Interpreting the *K* value

Recall that when 10^{-3} < K < 10^3, the reaction is considered to contain a significant amount of both reactants and products at equilibrium.

ANSWER:

- The equilibrium lies far to the right.
- The reaction will proceed very slowly.
- (•) The reaction contains significant amounts of products and reactants at equilibrium.
- The equilibrium lies far to the left.

Correct

Pause and Predict Video Quiz: Equilibrium and Equilibrium Constants

First, launch the video below. You will be asked to use your knowledge of chemistry to predict the outcome of an experiment. Then, close the video window and answer the questions at right. You can watch the video again at any point.

Part A

A chemical reaction is at equilibrium when

Hint 1. How to approach the problem

Watch the video to review the definition of chemical equilibrium.

ANSWER:

- the reaction occurs quickly.
- the rate of the forward reaction equals the rate of the reverse reaction.
- there is no reverse reaction possible, only a forward reaction.
- there are equal amounts of products and reactants.

Correct

Part B

A reaction that has a large equilibrium constant has

Hint 1. How to approach the problem

Recall that a reaction's equilibrium constant is the ratio of products to reactants when the reaction is at equilibrium.

ANSWER:

- more products then reactants at equilibrium.
- more reactants then products at equilibrium.
- equal amounts of products and reactants at equilibrium.
- a faster reaction rate than a reaction with a small equilibrium constant.

Correct

Part C

If a reaction is exothermic and produces large quantities of heat at equilibrium, its equilibrium constant will be

Hint 1. How to approach the problem

In an exothermic reaction, heat is considered a product. Think about how the equilibrium constant is calculated: ratio of products to reactants.

ANSWER:

- large.
- negative, because it is an exothermic reaction.
- small (less than 1).
- zero, because the reaction will not have an equilibrium constant.

Correct

Part D

Consider the reaction

$$\rm N_2O_4(g) \rightarrow NO_2(g)$$

where $\rm N_2O_4$ is colorless and $\rm NO_2$ is brown. At a temperature of $5\,\rm ^\circ C$, the reaction is colorless at equilibrium. This indicates that the equilibrium constant for this reaction at this temperature is

Hint 1. How to approach this problem

Think about the color of the reactants and the products and how the color of the reaction at equilibrium relates to the equilibrium constant.

ANSWER:

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- negative.
- zero, because equilibrium can only be obtained at room temperature.
- small (less than 1).
- large.

Correct

Visual Representations of Equilibrium

Chemical equilibrium is the reaction state when the concentrations of reactants and products remain constant over time. The reaction does not stop, but it reaches a dynamic equilibrium when the rates of the forward and reverse reactions become equal. A mixture of reactants and products in the equilibrium state is called an equilibrium mixture. Let's consider a general reversible reaction where \rm A and \rm B are the reactants and \rm C and \rm D are the products:

$$a{\rm A}+b {\rm B}\rightleftharpoons c{\rm C}+d {\rm D}$$

The concentrations in an equilibrium mixture are related by the equilibrium equation

$$\large{K_{\rm c}=\frac{[\rm C]^c[\rm D]^d}{[\rm A]^a[\rm B]^b}}$$

where \texttip{K_{\rm c}}{K_c} is the equilibrium constant.

Part A

The following pictures represent the equilibrium mixtures of five different chemical reactions. All the reactions proceed according to the general balanced chemical reaction

$$\rm A+B \rightleftharpoons AB$$

Red spheres represent \rm A atoms, blue spheres are \rm B atoms, and red-blue clusters are \rm AB molecules. Rank the reactions in order of decreasing equilibrium constant.

Rank the reactions from largest to smallest equilibrium constant To rank items as equivalent, overlap them.

Hint 1. How to approach the problem

First, use the the chemical equation to write the equilibrium-constant expression.
Next, relate the molecules we see on the pictures to the equilibrium molar concentrations in the \texttip{K_{\rm c}}{K_c} expression. Molarity is the number of moles divided by volume, and the number of moles is the number of molecules divided by Avogadro's number. Therefore, the molar concentrations in the \texttip{K_{\rm c}}{K_c} expression will be directly proportional to the number of molecules you see in the pictures.
Since we don't know the volume, we can't calculate the actual concentrations and \texttip{K_{\rm c}}{K_c} values, but we can compare them and rank them accordingly.

Hint 2. Select the correct equilibrium-constant expression

Which of these is the correct equilibrium-constant expression for the given reaction?

ANSWER:

- $K_{\rm c}=[\rm A][\rm B][\rm AB]$
- $\large{K_{\rm c}=\frac{[\rm A][\rm B]}{[\rm AB]}}$
- $\large{K_{\rm c}=\frac{[\rm A][\rm B]}{[\rm AB]^2}}$
- $\large{K_{\rm c}=\frac{[\rm AB]}{[\rm A][\rm B]}}$
- $\large{K_{\rm c}=\frac{[\rm AB]^2}{[\rm A][\rm B]}}$

ANSWER:

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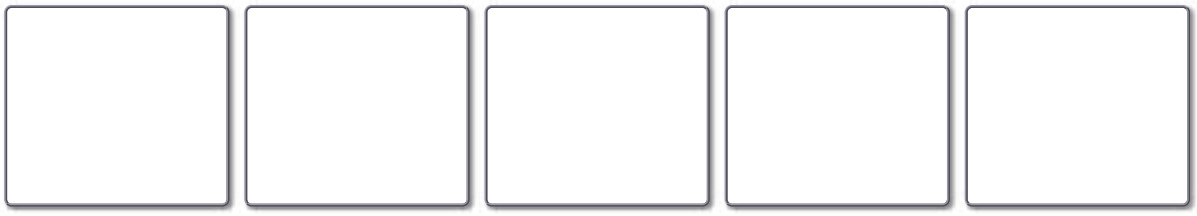
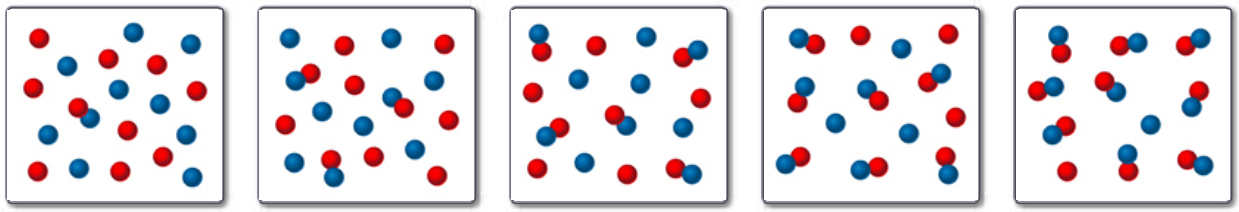
Reset Help



Smallest equilibrium constant Largest equilibrium constant

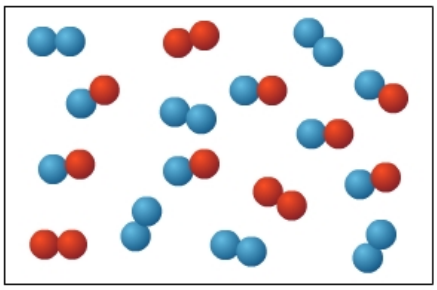


The correct ranking cannot be determined.

Correct

Part B

A chemical reaction between \rm X_2 (red) and \rm Y_2 (blue) produces \rm XY (red-blue). All compounds are in a gaseous state. The picture shown here represents the equilibrium mixture.



Calculate the equilibrium constant for this reaction.

Express your answer numerically.

Hint 1. How to approach the problem

First, write the balanced chemical equation. Use the balanced equation to write the equilibrium-constant expression.
Molarity is the number of moles divided by volume, and the number of moles is the number of molecules divided by Avogadro's number. Therefore, the molar concentrations in the \texttip{K_{\rm c}}{K_c} expression will be directly proportional to the number of molecules you see in the picture (three \rm X_2 molecules, six \rm Y_2 molecules, and seven \rm XY molecules).

Hint 2. Select the equation for the reaction

Which of the following is the correct balanced equation for the reaction?

ANSWER:

- \rm X_2+Y_2 \rightleftharpoons XY
- \rm X_2+Y_2 \rightleftharpoons 2XY
- \rm X_2+Y_2 \rightleftharpoons 4XY
- \rm 3X_2+6Y_2\rightleftharpoons7XY

Hint 3. Select the correct equilibrium-constant expression

Which of the following expressions is the correct form of the equilibrium constant for the given reaction?

ANSWER:

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- K_{\rm c}= [\rm X_2][\rm Y_2][\rm XY]
- \large{K_{\rm c}= \frac {[\rm X_2][\rm Y_2]}{[\rm XY]}}
- \large{K_{\rm c}= \frac {[\rm X_2][\rm Y_2]}{[\rm XY]^2}}
- \large{K_{\rm c}= \frac {[\rm XY]}{[\rm X_2][\rm Y_2]}}
- \large{K_{\rm c}= \frac {[\rm XY]^2}{[\rm X_2][\rm Y_2]}}

ANSWER:

\texttip{K_{\rm c}}{K_c} = 2.72

Correct

If we put the concentrations of the products in the denominator of the equilibrium-constant expression and the concentrations of the reactants in the numerator, we get the reciprocal expression that represents the equilibrium constant for the reverse reaction. K_{\rm c}'={1}/{K_{\rm c}} is the equilibrium constant for the reaction \rm 2XY \rightleftharpoons X_2+Y_2.

The Equilibrium-Constant Expression

Learning Goal:

To understand the form and meaning of the equilibrium-constant expression.

Chemical reactions are usually not one-way trips from reactant to product. In fact, most chemical reactions are reversible to at least some extent. Products formed by the forward reaction may react with each other to regenerate the reactants. When reactants are mixed, they will begin to react at a forward reaction rate particular to that chemical reaction. As reactants are depleted and products are formed, however, the rate of the forward reaction begins to slow, and the rate of the reverse reaction begins to increase.

Eventually, the forward and reverse reaction rates will be identical. Equilibrium is *dynamic*, meaning that it is a balance of continuous forward and reverse reactions. To indicate this dynamic nature of equilibrium, chemical equations of reactions with measurable reverse reaction rate are written using a double-headed arrow:

a{\rm A} + b{\rm B} \rightleftharpoons c{\rm C} + d{\rm D}

The equilibrium-constant expression

The equilibrium-constant expression is used to describe the concentration of reactants and products for a reaction in dynamic equilibrium. For ideal gases and ideal solutions in homogeneous equilibria, where all reactants and products are in the same phase, the extent to which a particular chemical reaction proceeds to products is given by the *equilibrium equation*

a{\rm A} + b{\rm B} \rightleftharpoons c{\rm C} + d{\rm D}, \large{~~~~~K = \frac {[{\rm C}]^c[{\rm D}]^d}{[{\rm A}]^a[{\rm B}]^b}}

where \texttip{K}{K} is the *equilibrium constant* and the right-hand side of the equation is known as the *equilibrium-constant expression*.

The concentration of each product raised to its coefficient is divided by the concentration of each reagent raised to its coefficient according the the balanced chemical equation. Therefore, the higher the concentration of products, the larger the value of \texttip{K}{K} will be.

Part A

Identify the proper form of the equilibrium-constant expression for the equation

\rm N_2(g) + O_2(g) \rightleftharpoons 2 NO(g)

Hint 1. How to approach the problem

Take a look at the chemical reaction equation. To write the chemical reaction equation, multiply the concentrations of each of the products. Be sure to raise each product by its stoichiometric coefficient. Then, divide by the concentrations of the reactant, with each reactant raised to the power of its stoichiometric coefficient.

ANSWER:

- \large{K=\rm \frac{[NO]}{[N_2][O_2]}}
- \large{K=\rm \frac{[NO]^2}{[N_2][O_2]}}
- \large{K=\rm \frac{[N_2][O_2]}{[NO]^2}}
- \large{K=\rm \frac{2[NO]}{[N_2][O_2]}}

Correct

Understanding the magnitude of \texttip{K}{K}

By analyzing the equilibrium-constant expression we can determine that a larger concentration of products in the numerator of the expression will result in larger values fo[illegible]er values of \texttip{K}{K} are associated with reactions running toward completion, whereas smaller values of \texttip{K}{K} are associated with a

larger concentration of reactants.

Value of K	Reaction favors	Reaction lies to
K \lll 1	reactants	left
K\sim 1	neither reactants nor products	center
K \ggg 1	products	right

Note that reaction stoichiometries with a large difference between the number of moles of reactants and products may not strictly follow these general guidelines.

Part B

The equilibrium-constant of the reaction

$$\rm NO_2(g) + NO_3(g) \rightleftharpoons N_2O_5(g)$$

is K = 2.1 \times 10^{-20}. What can be said about this reaction?

Hint 1. How to approach the problem

Consider how the numerical value of \texttip{K}{K} is influenced by the magnitude of the concentrations of the reactants and the products. For instance, if the concentrations of products are relatively large at equilibrium compared to the concentrations of reactants, the numerator of the equilibrium-constant expression will be quite large compared to the denominator, and the value of \texttip{K}{K} will be correspondingly large.

ANSWER:

- At equilibrium the concentration of products and reactants is about the same.
- At equilibrium the concentration of products is much greater than the concentration of reactants.
- At equilibrium the concentration of reactants is much greater than that of products.
- There are no reactants left over once the reaction reaches equilibrium.

Correct

Part C

Where does the equilibrium of this reaction lie?

Hint 1. How to determine where the reaction equilibrium lies

Recall that if products are favored, the equilibrium is said to lie to the right, since more of the chemical species from the right-hand side of the equation are at equilibrium. Similarly, if reactants are favored, the equilibrium is said to lie to the left, since more of the chemical species from the left-hand side of the equation will be present at equilibrium.

ANSWER:

- To the left
- To the right
- To neither the left nor the right

Correct

The equilibrium-constant expression for heterogeneous equilibria

Heterogeneous equilibria involve reactants and products that are not all in the same phase. For example, a solid may decompose, forming two gases. When writing equilibrium-constant expressions for heterogeneous reactions, generally the concentrations of pure solids and liquids are omitted. The composition of a pure solid or liquid does not change over the course of the reaction; only its quantity changes. Since its concentration isn't changing, the values can usually be excluded.

Given the equation x{\rm CD(s)} \rightleftharpoons c{\rm C(g)}+d{\rm D(g)}, the equilibrium constant expression would be

$$K=[{\rm C}]^c[{\rm D}]^d$$

Part D

The acid \rm HOCl (hypochlorous acid) is produced by bubbling chlorine gas through a suspension of solid mercury(II) oxide particles in liquid water according to the equation

$$\rm 2HgO(s) + H_2O(l) + 2Cl_2(g) \rightleftharpoons 2HOCl(aq) + HgO \cdot HgCl_2(s)$$

What is the equilibrium-constant expression for this reaction?

Hint 1. How to approach the problem

y the gases and aqueous species

Pure solids and liquids are not included in equilibrium-constant expressions. Only those species with concentration values, such as gases and aqueous solutions, will be included.

Select all of the gases and aqueous solutions present in the reaction

$$\rm 2HgO(s) + H_2O(l) + 2Cl_2(g) \rightleftharpoons 2HOCl(aq) + HgO \cdot HgCl_2(s)$$

Check all that apply.

ANSWER:

- [x] \rm HOCl
- [ ] \rm HgO \cdot HgCl_2
- [x] \rm Cl_2
- [ ] \rm HgO
- [ ] \rm H_2O

ANSWER:

- ○ \large{K=\rm \frac{[HOCl]^2[HgO \cdot HgCl_2]}{[Cl_2]^2[H_2O][HgO]^2}}
- ○ \large{K=\rm \frac{[HOCl]^2}{[Cl_2]^2[H_2O]}}
- ◉ \large{K=\rm \frac{[HOCl]^2}{[Cl_2]^2}}
- ○ \large{K=\rm \frac{[Cl_2]^2}{[HOCl]^2}}

Correct

± Calculating Equilibrium Constants

The equilibrium constant, \texttip{K}{K}, of a reaction at a particular temperature is determined by the concentrations or pressures of the reactants and products at equilibrium.

For a gaseous reaction with the general form

$$a{\rm A} + b{\rm B} \rightleftharpoons c{\rm C} + d{\rm D}$$

the \texttip{K_{\rm c}}{K_c} and \texttip{K_{\rm p}}{K_p} expressions are given by

$$\large{K_{\rm c} = \frac{[{\rm C}]^c[{\rm D}]^d}{[{\rm A}]^a[{\rm B}]^b}}$$

$$\large{K_{\rm p} = \frac{(P_{\rm C})^c(P_{\rm D})^d}{(P_{\rm A})^a(P_{\rm B})^b}}$$

The subscript \rm c or \rm p indicates whether \texttip{K}{K} is expressed in terms of concentrations or pressures. Equilibrium-constant expressions do not include a term for any pure solids or liquids that may be involved since their composition does not change throughout the reaction. The standard state of a pure substance is the pure substance itself, and although the quantity may change the sample remain pure. The concentration is effectively equal to 1, and will not impact the magnitude of \texttip{K}{K}.

Part A

Phosgene (carbonyl chloride), \rm COCl_2, is an extremely toxic gas that is used in manufacturing certain dyes and plastics. Phosgene can be produced by reacting carbon monoxide and chlorine gas at high temperatures:

$$\rm CO(g)+ Cl_2(g) \rightleftharpoons COCl_{2}(g)$$

Carbon monoxide and chlorine gas are allowed to react in a sealed vessel at 454 {\rm \, ^{\circ}\rm C} . At equilibrium, the concentrations were measured and the following results obtained:

Gas	Partial Pressure (\rm atm)
\rm CO	0.730
\rm Cl_2	1.18
\rm COCl_2	0.230

What is the equilibrium constant, \texttip{K_{\rm p}}{K_p}, of this reaction?

Express your answer numerically.

Hint 1. Identify the equilibrium-constant expression for the reaction

What is the expression for the equilibrium constant, \texttip{K_{\rm p}}{K_p}, for the reaction?

ANSWER:

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- \large{K_{\rm p} = \frac{P_{\rm CO} \cdot P_{\rm Cl_2}}{P_{\rm COCl_2}}}
- \large{K_{\rm p} = \frac{P_{\rm COCl_2}}{P_{\rm CO} \cdot P_{\rm Cl_2}}}
- \large{K_{\rm p} = \frac{1}{P_{\rm CO} \cdot P_{\rm Cl_2}}}
- \large{K_{\rm p} = \frac{(P_{\rm COCl_2})^{2}}{P_{\rm CO} \cdot (P_{\rm Cl_2})^{2}}}

ANSWER:

\texttip{K_{\rm p}}{K_p} = 0.267

Correct

Deriving concentrations from data

In Part A, you were given the equilibrium pressures, which could be plugged directly into the formula for \texttip{K}{K}. In Part B however, you will be given initial concentrations and only one equilibrium concentration. You must use this data to find all three equilibrium concentrations before you can apply the formula for \texttip{K}{K}.

Part B

The following reaction was performed in a sealed vessel at 767 {\rm \, ^{\circ}\rm C} :

$$\rm H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)$$

Initially, only \rm H_2 and \rm I_2 were present at concentrations of \rm [H_2]= 3.65{\rm \it M} and \rm [I_2]= 2.15{\rm \it M}. The equilibrium concentration of \rm I_2 is 0.0900 {\rm \it M} . What is the equilibrium constant, \texttip{K_{\rm c}}{K_c}, for the reaction at this temperature?

Express your answer numerically.

Hint 1. Identify the equilibrium-constant expression

What is the equilibrium-constant expression for this reaction?

ANSWER:

- \large{K_{\rm c} = \rm \frac{[H_2][I_2]}{[HI]^2}}
- \large{K_{\rm c} = \rm \frac{[HI]}{[H_2][I_2]}}
- \large{K_{\rm c} = \rm \frac{[HI]}{[H_2]^2[I_2]^2}}
- \large{K_{\rm c} = \rm \frac{[HI]^2}{[H_2][I_2]}}

Hint 2. Determine the change in the concentration of iodine

What value should be placed in the table in place of the "?" for the change in concentration of the reactant \rm I_2?

$$\begin{array}{l} \rm Concentration~({\it M}) & \rm H_2 & + & \rm I_2 & \rightleftharpoons & \rm 2HI \\ \rm Initial & 3.65 & & 2.15 & & 0 \\ \rm Change & ? & & ? & & \\ \rm Equilibrium & & 0.0900 & & & \end{array}$$

Be sure to include the *sign* of the change.

Express your answer with the appropriate units.

ANSWER:

-2.06 {\rm M}

Hint 3. Determine the change in the concentration of hydrogen

What value should be placed in the table in place of the "?" for the change in concentration of the reactant \rm H_2?

$$\begin{array}{l} \rm Concentration~({\it M}) & \rm H_2 & + & \rm I_2 & \rightleftharpoons & \rm 2HI \\ \rm Initial & 3.65 & & 2.15 & & 0 \\ \rm Change & ? & & -2.06 & & \\ \rm Equilibrium & & 0.0900 & & & \end{array}$$

Be sure to include the *sign* of the change.

Express your answer with the appropriate units.

ANSWER:

-2.06 {\rm M}

Hint 4. Determine the change in the concentration of hydrogen iodide

What value should be placed in the table in place of the "?" for the change in concentration of the product \rm HI?

$$\begin{array}{l} \rm Concentration~({\it M})&\rm H_2&+&\rm I_2&\rightleftharpoons&\rm 2HI\cr \rm Initial&3.65& &2.15& &\rm 0 \cr \rm Change&-2.06 &&-2.06&& ? \cr \rm Equilibrium& && 0.0900&&& && &\end{array}$$

Express your answer with the appropriate units.

ANSWER:

4.12 {\rm M}

Hint 5. Determine the equilibrium concentration of hydrogen

What value should be placed in the table in place of the "?" for the equilibrium concentration of the reactant \rm H_2?

$$\begin{array}{l} \rm Concentration~({\it M})&\rm H_2&+&\rm I_2&\rightleftharpoons&\rm 2HI\cr \rm Initial&3.65& &2.15& &\rm 0 \cr \rm Change&-2.06 &&-2.06 && 4.12 \cr \rm Equilibrium&? && 0.0900&&& && &\end{array}$$

Express your answer with the appropriate units.

ANSWER:

\rm [H_2] = 1.59 {\rm M}

Hint 6. Determine the equilibrium concentration of hydrogen iodide

What value should be placed in the table in place of the "?" for the equilibrium concentration of the reactant \rm HI?

$$\begin{array}{l} \rm Concentration~({\it M})&\rm H_2&+&\rm I_2&\rightleftharpoons&\rm 2HI\cr \rm Initial&3.65& &2.15& &\rm 0 \cr \rm Change&-2.06 &&-2.06 && 4.12 \cr \rm Equilibrium&1.59 && 0.0900&&?\end{array}$$

Express your answer with the appropriate units.

ANSWER:

\rm [HI] = 4.12 {\rm M}

ANSWER:

\texttip{K_{\rm c}}{K_c} = 119

Correct

Analysis of Equilibria

Learning Goal:

To understand how reactions at equilibrium respond to changes in conditions, including volume changes and addition of reactants or products.

Equilibrium equations can be used to determine what will happen to a system at equilibrium if certain conditions are changed (e.g., the addition of a reactant or product or—in the case of gaseous reactions—a volume change).
For example, given the reaction

$$\large{\rm \frac{1}{8}S_8(s)\rm + O_2(g)\rm \rightleftharpoons SO_2(g)}$$

what will happen to the system if more \rm SO_2(g) is added so that the concentration of \rm SO_2(g) is doubled?

The equilibrium equation is given by

$$\large{K = \rm \frac{[SO_2]}{[O_2]}}$$

Equilibrium equations do not include expressions for any pure solids or liquids that may be involved in the reaction. This is because at any given temperature their concentration is constant, and these values are by convention included in the value of the equilibrium constant.

If the system is perturbed in such a manner that when equilibrium is reestablished if the value of \rm [SO_2] is doubled, then the value of \rm [O_2] must also be double since the value of \texttip{K}{K} remains constant at any given temperature.

Part A

For the decomposition of \rm A to \rm B and \rm C,

$$\rm A(s) \rightleftharpoons B(g) + C(g)$$

how will the reaction respond to each of the following changes at equilibrium?

Drag the appropriate items to their respective bins.

Hint 1. Determine how each change affects \texttip{Q}{Q}

At equilibrium, \texttip{Q}{Q} is equal to \texttip{K}{K}. What happens to \texttip{Q}{Q} as a result of each of the following changes?

Drag the appropriate items to their respective bins.

Hint 2. Determine the expression for the reaction quotient

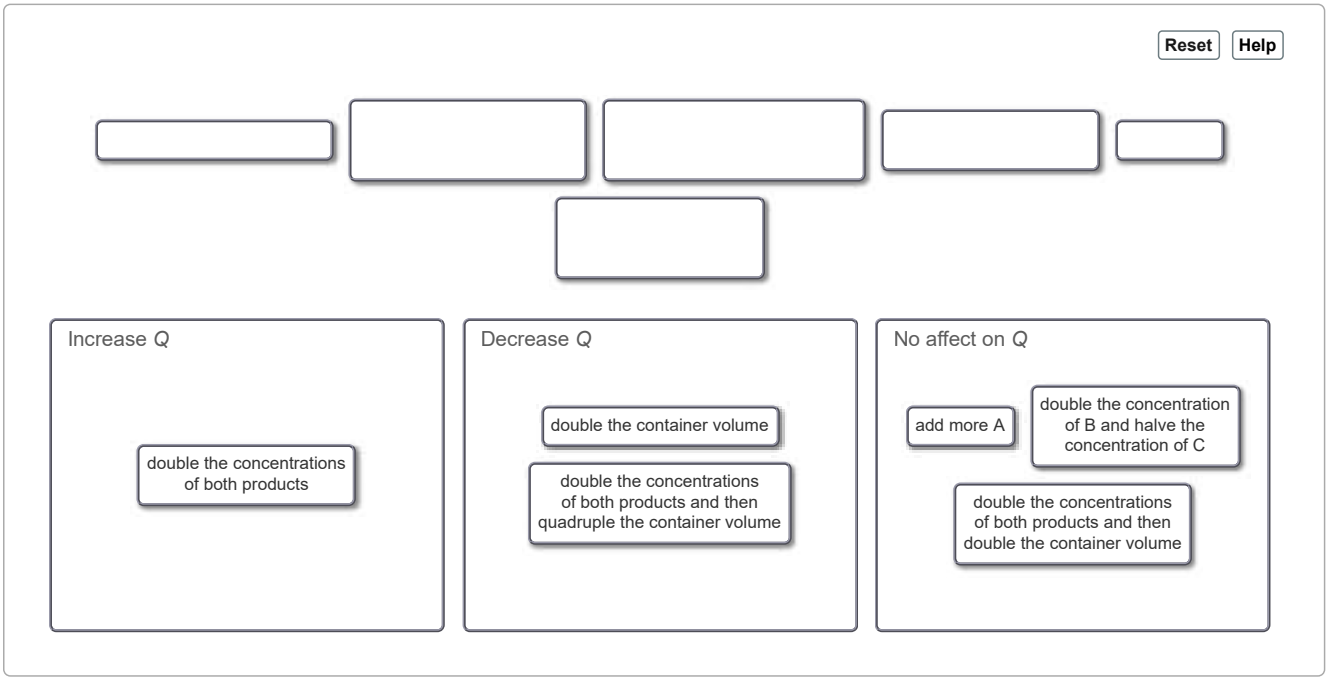
What is the expression for the reaction quotient \texttip{Q}{Q} of the reaction?

Express your answer in terms of \rm [A], \rm [B], and \rm [C].

ANSWER:

\texttip{Q}{Q} = \left[B\right] \left[C\right]

ANSWER:

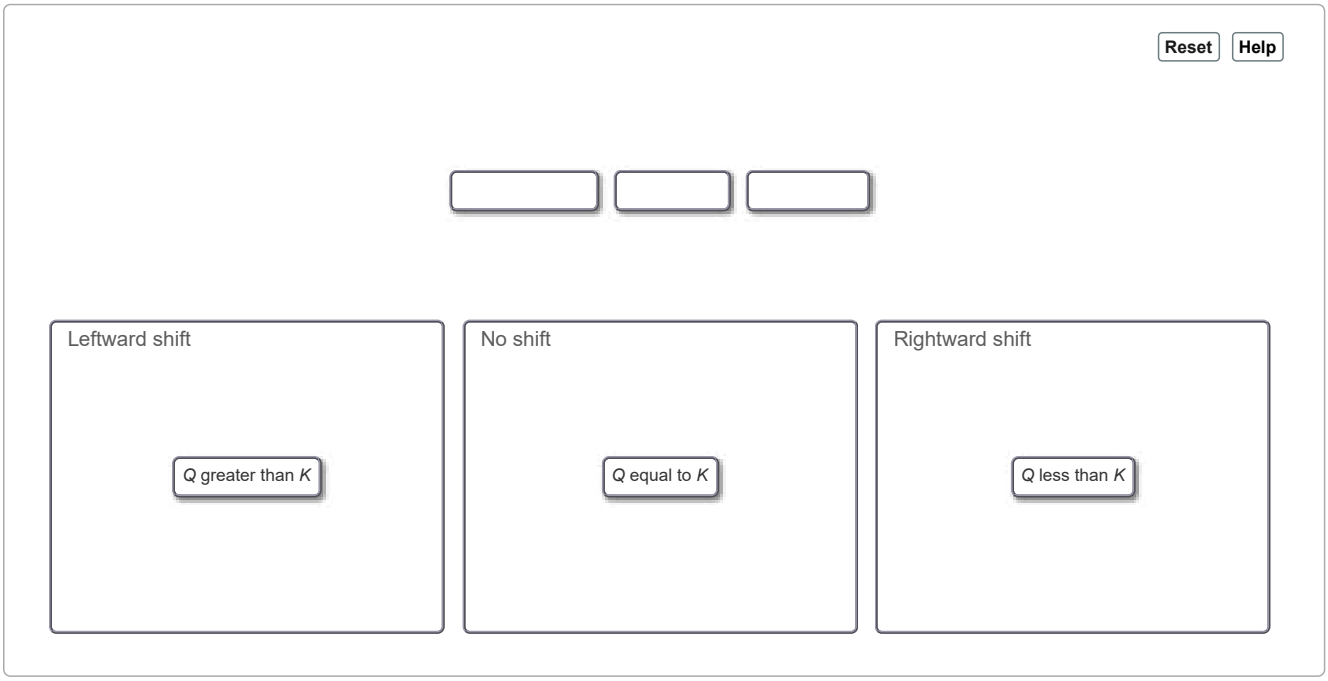


Hint 2. Determine how a reaction responds during different conditions

Identify how a reaction responds to each of these conditions.

Drag the appropriate items to their respective bins.

ANSWER:



ANSWER:

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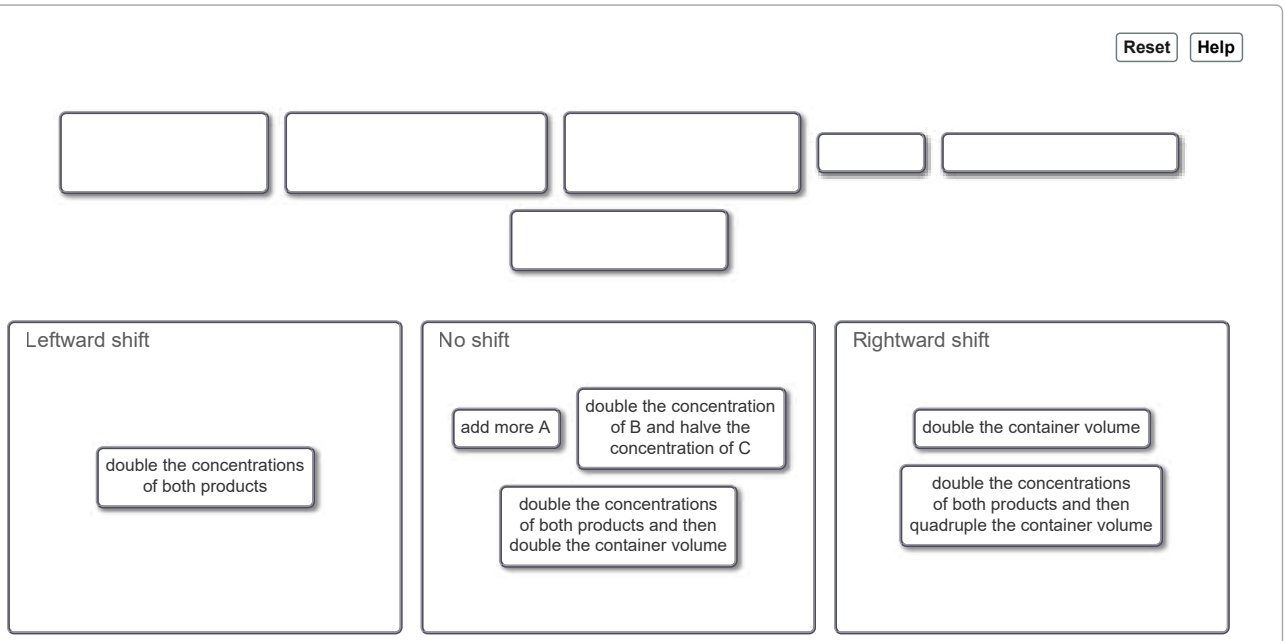


Correct

Part B

For the reaction of \rm A and \rm B forming \rm C,

$$\rm A(g)+B(s) \rightleftharpoons 2C(g)$$

how will the reaction respond to each of the following changes at equilibrium?

Drag the appropriate items to their respective bins.

Hint 1. Determine how each change affects \texttip{Q}{Q}

At equilibrium, \texttip{Q}{Q} is equal to \texttip{K}{K}. What happens to \texttip{Q}{Q} as a result of each of the following changes?

Drag the appropriate items to their respective bins.

Hint 1. Determine the expression for the reaction quotient

What is the expression for the reaction quotient \texttip{Q}{Q} of the reaction?

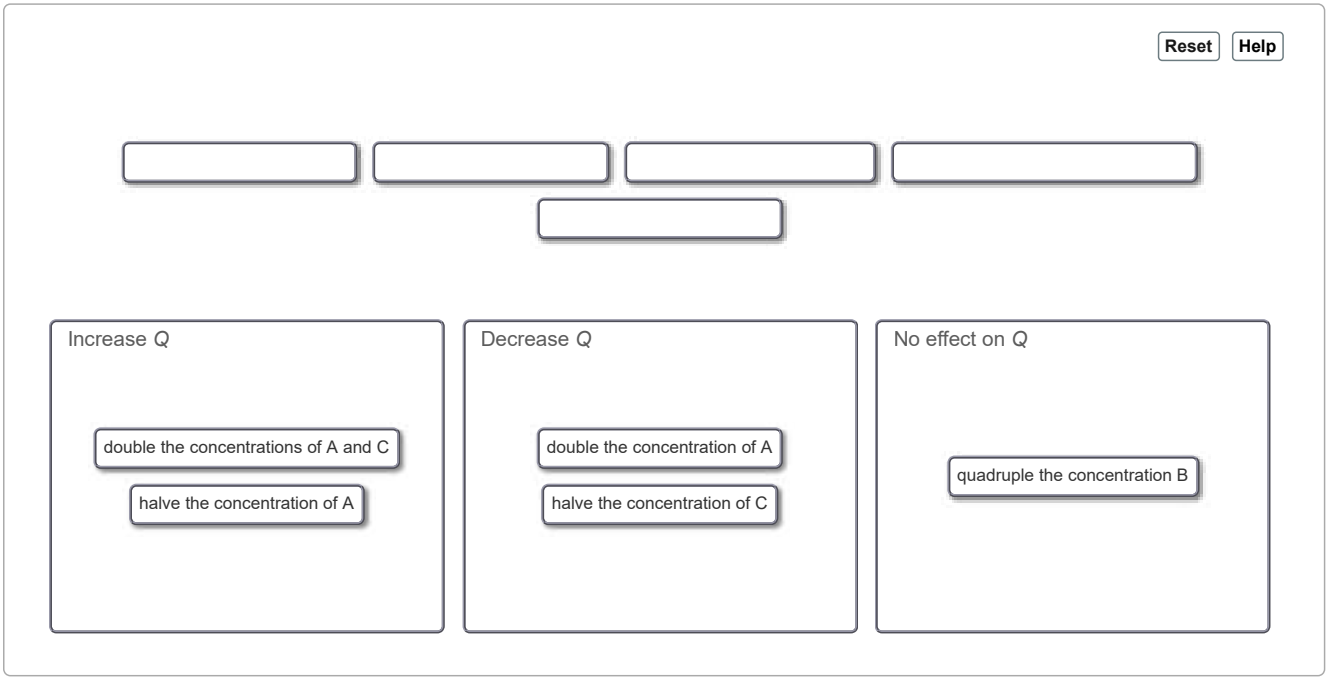
Express your answer in terms of \rm [A], \rm [B], and \rm [C].

ANSWER:

\texttip{Q}{Q} = \large{\frac{\left[C\right]^{2}}{\left[A\right]}}

ANSWER:

Typesetting math: 68%

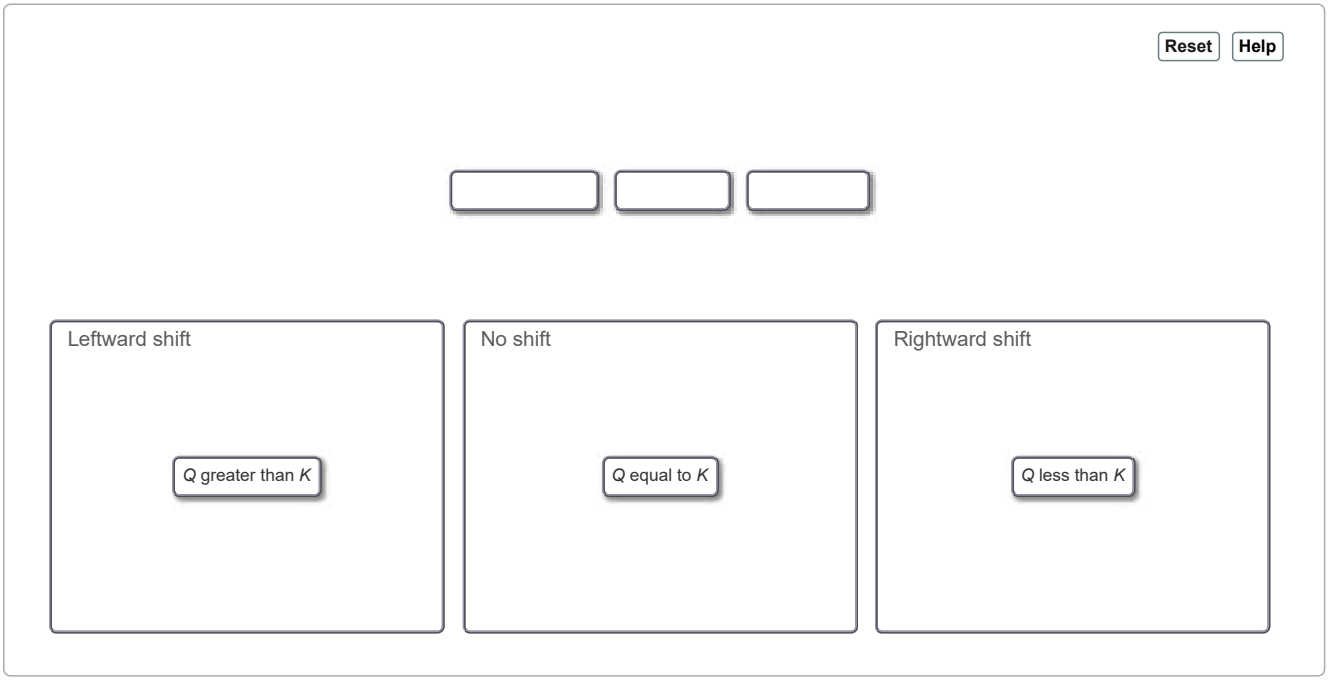


Hint 2. Determine how a reaction responds during different conditions

Identify how a reaction responds to each of these conditions.

Drag the appropriate items to their respective bins.

ANSWER:



ANSWER:

Reset Help

Leftward shift

halve the concentration of A

double the concentrations of A and C

No shift

quadruple the concentration of B

Rightward shift

halve the concentration of C

double the concentration of A

Correct

Applying Le Châtelier's Principle

When a chemical reaction is at equilibrium, \texttip{Q}{Q} (the reaction quotient) is equal to \texttip{K}{K} (the equilibrium constant). If a stress is applied to the mixture that changes the value of \texttip{Q}{Q}, then the system is no longer at equilibrium. To regain equilibrium, the reaction will either proceed forward or in reverse until \texttip{Q}{Q} is equal to \texttip{K}{K} once again. Alternatively, equilibrium can be disrupted by a change in temperature, which changes the value of \texttip{K}{K}. The result however is the same, and the reaction will proceed forward or in reverse until \texttip{Q}{Q} is equal to the new \texttip{K}{K}.

Le Châtelier's principle summarizes this idea:
If a stress is applied to a reaction mixture at equilibrium, a net reaction occurs in the direction that relieves the stress.

Part A

Consider the following system at equilibrium:

$$\rm 3A(aq) + 3B(aq) \rightleftharpoons 5C(aq)$$

Classify each of the following actions by whether it causes a leftward shift, a rightward shift, or no shift in the direction of the net reaction.

Drag the appropriate items to their respective bins.

Hint 1. How to approach the problem

One way to approach the problem is consider each situation qualitatively and determine which shift will "undo" the effects of the stress.
Another approach is to look at each situation quantitatively. For example, calculate \texttip{K}{K} if \rm A, \rm B, and \rm C are each 1 M. Then calculate \texttip{Q}{Q} with increased or decreased values as the choices indicate.

- If Q>K, the reaction will shift toward the reactants.
- If Q<K, the reaction will shift toward the products.

Hint 2. Identify the expression for the equilibrium constant

Which of the following expressions correctly describes \texttip{K_{\rm c}}{K_c} for the reaction?

ANSWER:

- \large{\rm {\it K}_c = \frac{[C]}{[A]^3[B]^3}}
- \large{\rm {\it K}_c = \frac{[C]^5}{[A]^3[B]^3}}
- \large{\rm {\it K}_c = \frac{[C]^5}{[A][B]^3}}
- \large{\rm {\it K}_c = \frac{[C]^5}{[A]^3[B]}}

Hint 3. Identify how an increase in reactant concentration affects the reaction

If the concentration of a reactant increases, how will the reaction respond to relieve the stress?

ANSWER:

Typesetting math: 68%

- It will consume reactants to make products.
- It will consume products to make reactants.

Hint 4. Identify how a decrease in reactant concentration affects the reaction

If the concentration of a reactant decreases, how will the reaction respond to relieve the stress?

ANSWER:

- It will consume reactants to make products.
- It will consume products to make reactants.

Hint 5. Identify how an increase in product concentration affects the reaction

If the concentration of the product increases, how will the reaction respond to relieve the stress?

ANSWER:

- It will consume reactants to make products.
- It will consume products to make reactants.

Hint 6. Identify how a decrease in product concentration affects the reaction

If the concentration of the product decreases, how will the reaction respond to relieve the stress?

ANSWER:

- It will consume reactants to make products.
- It will consume products to make reactants.

ANSWER:

Reset Help

Leftward shift
- Double both [B] and [C]
- Increase [C]
- Decrease [A]
- Decrease [B]

Rightward shift
- Increase [A]
- Increase [B]
- Decrease [C]

No shift
- Double [A] and reduce [B] to one half

Correct

Part B

The following system is at equilibrium:

$$\rm 3M(g) + N(s) \rightleftharpoons 2L(g)$$

Classify each of the following actions by whether it causes a leftward shift, a rightward shift, or no shift in the direction of the net reaction.

Drag the appropriate items to their respective bins.

Hint 1. How to approach the problem

In a gas reaction, it is useful to look at the number of moles of gaseous reactant versus the number of moles of gaseous product. In this case, there are three moles of gaseous reactant, \rm 3 M(g), and two moles of gaseous product, \rm 2 L(g).

The result of an increase in volume is a decrease in pressure. This stress can be relieved if the reaction moves toward the side that exerts more pressure (i.e., has more moles of gas). Similarly, the result of a decrease in volume is an increase in pressure. This stress can be relieved if the reaction moves toward the side that exerts less pressure (i.e., has fewer moles of gas).

Hint 2. Identify the expression for the equilibrium constant

Which of the following expressions correctly describes K_{\rm p} for the reaction?

ANSWER:

- \large{\rm {\it K}_p = \frac{({\it P}_L)^2}{({\it P}_M)({\it P}_N)}}
- \large{\rm {\it K}_p = \frac{({\it P}_L)^2}{({\it P}_M)^3({\it P}_N)}}
- \large{\rm {\it K}_p = \frac{({\it P}_L)^2}{({\it P}_M)^3}}
- \large{\rm {\it K}_p = \frac{({\it P}_L)^2}{({\it P}_M)}}

Hint 3. Identify how a decrease in volume affects the reaction

If the volume is decreased, how will the reaction respond to relieve the stress?

ANSWER:

- It will consume reactants to make products.
- It will consume products to make reactants.

Hint 4. Identify how an increase in volume affects the reaction

If the volume is increased, how will the reaction respond to relieve the stress?

ANSWER:

- It will consume reactants to make products.
- It will consume products to make reactants.

ANSWER:

Reset Help

Leftward shift	Rightward shift	No shift
Increase the volume	Decrease the volume	Add more N Remove some N

Correct

For a certain chemical reaction, \Delta H^{\circ}= -156\; \rm kJ. Assuming the reaction is at equilibrium, classify each of the following actions by whether it causes a leftward shift, a rightward shift, or no shift in the direction of the net reaction.

Drag the appropriate items to their respective bins.

Hint 1. How to approach the problem

An endothermic reaction absorbs heat:

$$\rm heat + reactants \rightleftharpoons products$$

By thinking of heat as a reactant, we can treat a temperature change the same way we would a change in reactant concentration.

In contrast, an exothermic reaction releases heat:

$$\rm reactants \rightleftharpoons products+heat$$

By thinking of heat as a product, we can treat a temperature change the same way we would a change in product concentration.

Hint 2. Identify the correct balanced equation that includes the heat contribution

Which of the following balanced equations correctly shows the contribution of the heat in this reaction?

ANSWER:

- \rm heat + reactants \rightleftharpoons products
- \rm reactants \rightleftharpoons products + heat

Hint 3. Identify how an increase in temperature affects the reaction

If the temperature is increased, how will the reaction respond to relieve the stress?

ANSWER:

- It will consume reactants to make products.
- It will consume products to make reactants.

Hint 4. Identify how a decrease in temperature affects the reaction

If the temperature is decreased, how will the reaction respond to relieve the stress?

ANSWER:

- It will consume reactants to make products.
- It will consume products to make reactants.

ANSWER:

Reset Help

Leftward shift

Increase the temperature

Rightward shift

Decrease the temperature

No shift

Typesetting math: 68%

Correct

For an exothermic reaction, a temperature increase causes the value of \texttip{K}{K} to decrease. Then, because Q>K, the reaction shifts toward reactants. The opposite occurs for endothermic reactions.

Score Summary:

Your score on this assignment is 126%.
You received 126.18 out of a possible total of 100 points.