**Name: \_\_\_\_\_\_\_\_\_\_\_\_ Date:**

**STUDENT ACTIVITY - Chemical Equilibrium**

***Introduction***

* Chemical equilibrium occurs when opposing reactions are proceeding at equal rates.
* The rate at which the products are formed from the reactants equals the rate at which the

reactants are formed from the products.

* As a result, concentrations cease to change, making the reaction appear to be stopped.
* How fast a reaction reaches this equilibrium state is a matter of kinetics.
* An ***equilibrium state*** results when a reaction is ***reversible.***

*kf*

*kr*

 **aA + bB cC + dD**

* At equilibrium the concentrations of reactants and products is still changing, however,

 ***the rate of the forward reaction (kf) is equal to the rate of the reverse reaction (kr)*** in what is called a ***dynamic equilibrium*** such that no change in their concentrations is observed. Thus, for equilibrium to occur, neither reactants nor products can escape from the system.

* The ***law of mass action*** states the ratio of forward and reverse processes is described by

 the ***equilibrium constant, Kc* ,** which can be calculated using a knowledge of the

 equilibrium concentrations of reactants and products.

* The equilibrium constant expression depends only on the stoichiometry of the reaction.

***Objectives and Success Criteria***

* Mastering the application of the ICE table methodology to equilibrium problems.
* Accurate solutions to problems involving reactant and product concentrations and

 equilibrium constants.

***MODEL 1:* The ICE Table**

*Example:*

Initially 1.50 moles of N2(g) and 3.50 moles of H2(g) were added to a 1 L container at 700 °C. As

a result of the reaction:

 **N2(g) + 3 H2(g) 2 NH3(g)**

the equilibrium concentration of NH3(g) became 0.540 M. What is the value of the equilibrium

constant for this reaction at the given temperature of 700 °C?

 **N2(g) + 3 H2(g) 2 NH3(g)**

**I.** Write the **I**nitial concentrations of

 reactants and products. **1.50 M 3.50 M 0 M**

**C**. Write the **C**hange in concentration due **−x −3x +2x**

 to reaction using the given reaction

 stoichiometric coefficients.

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**E**. Write the reactant and product **1.50 -x 3.50 -3x 2x**

 concentrations at **E**quilibrium.

Since we know that the equilibrium concentration of NH3(g) became 0.540 M, we can substitute that number for **2x**. Solving for **x**, we get **x = 0.270 M**. We can substitute these new numbers into the ICE Box chart.

 **N2(g) + 3 H2(g) 2 NH3(g)**

**I.** **1.50 M 3.50 M 0 M**

**C**. **− 0.270 − 0.810 + 0.540** \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**E**. **1.23 M 2.69 M 0.540 M**



We can now solve for the equilibrium constant *Kc* using the *equilibrium equation*:

 *equation. 1*

 *Kc* = [0.540]2\_\_\_\_\_ = **0.0122**

 [1.23] x [2.69]3

***Key Questions***

(i) In the above reaction we can monitor the change in concentration of reactants over time

and we can plot the data as follows:

 Label below

C

O

N

C

E

N

T

R

A

T

I

O

N

 (M)

 3 -

 \_\_\_\_\_\_\_\_\_\_

 2 -

 \_\_\_\_\_\_\_\_\_\_

 1 1 - \_\_\_\_\_\_\_\_\_\_

 a b c d

 time

1. Using the worked example from Model 1, label each data plot as either [H2], [N2] or [NH3].

2. Why does one data plot show an initial positive slope whereas the other two data plots show

 initial negative slopes?

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

3. Why does the uppermost plot have a steeper initial slope than the middle plot.

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4. At which time a, b, c or d is an equilibrium state first reached? \_\_\_\_\_\_\_\_

5. The equilibrium constant for this reaction was calculated to be 0.0122. What is *K*c for

 the reverse reaction? (Show work below).

6. Initially, 1.0 mol of NO(g) and 1 mol of Cl2(g) were added to a 1 L container. As a result

 of the reaction the equilibrium concentration of NOCl(g) became 0.96 M. Using the ICE Box

 methodology, determine the value of the equilibrium constant *K*c for this reaction:

 **2 NO(g) + Cl2(g) 2 NOCl(g)**

 I. \_\_\_\_\_\_\_M \_\_\_\_\_\_\_M \_\_\_\_\_\_\_M

 C. \_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

 E. \_\_\_\_\_\_\_M \_\_\_\_\_\_\_M \_\_\_\_\_\_\_M

Solve for *K*c

7. If a 10.00 L flask at 500 K is filled with a 0.30 mole of hydrogen and 0.30 mole of iodine.

 What are the equilibrium concentrations of all three gases?

 The equilibrium constant *K*c = 45.0.

 The relevant reaction is: **H2(g) + I2(g) 2 HI(g)**

:

 **[H2] = \_\_\_\_\_\_\_\_M [I2] = \_\_\_\_\_\_\_\_M [HI] = \_\_\_\_\_\_\_\_M**

8. Gaseous carbon dioxide is partially decomposed according to the equation below.

 An initial pressure of 1.00 atm of CO2 is placed in a closed container at 2500 K, and 2.10 %

 of the molecules decompose. Determine the equilibrium constant *K*p at this temperature.

 (*Kp* is solved the same way as *Kc*, except you will use pressure values instead of molarities).

 **2CO2(g) 2 CO(g) + O2(g)**

 ***Kp* = \_\_\_\_\_\_\_\_\_\_\_\_**