

Name:

Period:

Seat#:

Directions:

- Answer these on binder paper!
- The first couple problems are “chunked” for you to help you think through the steps.
- For ICE Table Problems, show your ICE Table and any algebra.
- If assuming the 5% rule, show that you checked that it was a valid assumption at the end.
 - Remember can only use 5% rule if $K < 1$
 - And usually $K = 1000x$ smaller than initial []’s is a better guess of when to use it
 - If $\frac{x}{\text{initial []}} \times 100 < 5\%$ then it is a valid assumption
- If 5% rule turns out not valid, then show your algebra or quadratic equation calculation
 - For quadratic equation make sure to indicate which answer for x is valid.
- For all other questions show calculations or give explanations when appropriate.
- Some answers are provided at the end. They are underlined.

1) For the reaction, $A \leftrightarrow 2B$, $K_c = 2$.

Suppose 3.0 moles of B and 3.0 moles of A are introduced into a 2.00 L flask.

- Calculate the [A] and [B]
- Is this system at equilibrium? Justify with showing a calculation.
- In which direction will the reaction proceed to reach equilibrium?
- As the system moves towards equilibrium what happens to the []’s of each chemical, increase or decrease?

2) $N_2(g) + O_2(g) \leftrightarrow 2 NO(g)$ The initial $[N_2] = 0.80 M$ and the initial $[O_2] = 0.20 M$ $K_c = 1.0 \times 10^{-5}$

- Based on the information given, and not doing any calculation, which direction does the reaction have to proceed to reach equilibrium? Explain your answer.
- Looking at the size of the K_c value, and the initial concentrations, do you predict that the 5% rule will be a valid assumption? Show why you think yes or no.
- Calculate the equilibrium concentrations for the reaction. $[N_2] = 0.8M, [O_2] = 0.2M, [NO] = 1.26E-3M$

3) $2NO_2 \leftrightarrow 2NO + O_2$ If 0.50 mol of NO_2 is placed in a 2.0L flask to create NO and O_2 $K_{eq} = 1.2 \times 10^{-5}$

- Calculate the $[NO_n]$
- Which way will the reaction proceed, right or left?
- As the system moves towards equilibrium what happens to the []’s of each chemical, increase or decrease?
- Looking at the size of the K_c value, and the initial concentrations, do you predict that the 5% rule will be a valid assumption? Show why you think yes or no.
- Calculate all the concentrations of each chemical once it reaches equilibrium.
 $[NO_2] = 0.25M, [NO] = 0.0114M, [O_2] = 0.00572M$

4) Calculate the equilibrium concentrations of all species if 3.000 moles of H_2 and 6.000 moles of F_2 are placed in a 3.000 L container. $H_2(g) + F_2(g) \leftrightarrow 2HF(g)$, $K_c = 1.15 \times 10^{-3}$ $[H_2] = 1M, [F_2] = 2M, [HF] = 0.048M$

5) At $650^\circ C$, the reaction below has a K_{eq} value of 0.771. If 2.00 mol of both hydrogen and carbon dioxide are placed in a 4.00 L container and allowed to react, what will be the equilibrium concentrations of all four gases? $[H_2] = 0.266M, [CO_2] = 0.266M, [CO] = 0.234M, [H_2O] = 0.234M$



Dougherty Valley HS Chemistry

Equilibrium – ICE Tables

- 6) $2\text{HI} \leftrightarrow \text{H}_2 + \text{I}_2$ $K_{\text{eq}} = 0.016$. The system starts with 0.010 M H_2 and I_2 and 0.096 M of HI
- Is this system at equilibrium? Justify with showing a calculation.
 - In which direction will the reaction proceed to reach equilibrium?
 - As the system moves towards equilibrium what happens to the []'s of each chemical, increase or decrease
 - Calculate all the concentrations of each chemical once it reaches equilibrium.
 $[\text{HI}] = 0.0644\text{M}$, $[\text{H}_2] = 0.0258\text{M}$, $[\text{I}_2] = 0.0258\text{M}$
- 7) 3.00 moles of N_2 gas and 1.00 mole of H_2 gas are combined in a 1 L reaction vessel. At equilibrium 0.663 moles of H_2 remain. $\text{N}_2 + 3\text{H}_2 \leftrightarrow 2\text{NH}_3$
- What are the resulting concentrations? $[\text{N}_2] = 2.89\text{M}$, $[\text{H}_2] = 0.663\text{M}$, $[\text{NH}_3] = 0.224\text{M}$
 - What is the value of the equilibrium constant at this particular temperature? $K_{\text{eq}} = .0596$
- 8) *Careful! There is something tricky about this problem!* ☺ Find the equilibrium constant, K_{eq} , for the following equilibrium situation. The initial concentrations of AB and A_2D are 0.30 M before they are mixed and when equilibrium is reached, the equilibrium concentration of A_2D is 0.20 M. $K_{\text{eq}} = 1.25$
- $$2\text{AB}(\text{g}) + \text{C}_2\text{D}(\text{s}) \leftrightarrow \text{A}_2\text{D}(\text{g}) + 2\text{CB}(\text{s})$$
- 9) At a particular temperature, Phosphorus pentachloride decomposes into Phosphorous trichloride and Chlorine gas. 0.500 moles of pure Phosphorus pentachloride is placed in a 2.00 L bottle and 0.7 M is the concentration of the chlorine gas in the same bottle. What are the resulting concentrations if the equilibrium constant at this particular temperature is $K_{\text{c}} = 6.5 \times 10^{-4}$? $[\text{PCl}_5] = 0.25\text{M}$, $[\text{PCl}_3] = 2.32\text{E-}4\text{M}$, $[\text{Cl}_2] = 0.7\text{M}$
- 10) ***NOT REQUIRED*** - *an extra one that has a quadratic equation in case you feel like you want to practice this kind. If you complete this and get it right I will give you some tickets* ☺ $2\text{HI} \leftrightarrow \text{H}_2 + \text{I}_2$ If $K_{\text{p}} = 50.5$ and the initial pressures are $\text{HI} = 0.975$ atm, $\text{H}_2 = 0.105$ atm and $\text{I}_2 = 0.215$ atm, what are the equilibrium pressures for all the substances?