EQUILIBRIUM appears around 4/75 multiple choice questions and in the free response every year.

-K_{eq} does not have units.
-Coefficients in balanced equations become exponents.
-Products go on top, Reactants on the bottom.
-PURE SOLIDS AND LIQUIDS ARE NOT INCLUDED

VOCABULARY

Equilibrium- a reversible chemical equation where reactants go to products and vise versa
Law of Mass Action-The relationship of the concentration of such products and reactants is given by the equilibrium expression

Molar Concentration

K_c –the constant used for molar concentrations

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

$$aA + bB \Leftrightarrow cC + dD$$

reactants products

Sample Problem:

Suppose you are given the following equilibrium:

$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$ Keq = 23.2 at 600 K

If the initial amounts of CO and H2O were both 0.100 M, what will be the amounts of each reactant and product at equilibrium?

Because you are only given the initials you assume that you have 0M for each product

	$CO(g)$	$+ H_2O$	$\rightleftharpoons CO_2(g)$	$+ H_2(g)$
Initial	0.100 M	0.100 M	0	0
Change				
Equilibrium				

Then some CO and H_2O will react meaning they are used up or subtracted as the same number of moles of the products are formed. This unknown value is noted as x.

	$CO(g)$	$+ H_2O$	$\rightleftharpoons CO_2(g)$	$+ H_2(g)$
Initial	0.100 M	0.100 M	0	0
Change	- x	-x	x	x
Equilibrium				

Then you subtract and add it all up at the bottom leaving you with unknown concentrations for each substance. Then from here you write the K_{eq} expression.

	$CO(g)$	$+ H_2O$	$\rightleftharpoons CO_2(g)$	$+ H_2(g)$
Initial	0.100 M	0.100 M	0	0
Change	- x	-x	x	x
Equilibrium	0.100 -x	0.100 -x	x	x

$$K_{eq} = \frac{[CO_2][H_2]}{[CO][H_2O]} = \frac{(x)(x)}{(0.100 - x)(0.100 - x)} = \frac{x^2}{(0.100 - x)^2} = 23.2$$

From here input both $x^2 / (0.100 - x)^2$ and 23.2 into the "y=" in a graphing calculator and calculate the intersect finding a value of x=0.0829

But because CO and H2O have a final value of 0.100 – x, you subtract it to get 0.017 M for the reactants and 0.0829M for the products

Pressure

K_p – The constant for partial pressure
-Solved in the same way BUT with
Pressures instead of gas.

$$K_P = \frac{P_C P_D}{P_A P_B}$$

$$K_P = K_c(RT)^{\Delta n}$$

-This equation is used to relate K_p with K_c.
- R = ideal gas constant .0821(L-atm)/(mol-K)
-T = absolute temperature (K)
- Δn = moles of product- moles of reactant

Le Chatelier's Principle

-States that whenever stress is placed on equilibrium, the equilibrium will shift to relieve the stress

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ ΔH= -92.6

	Right	Left
Concentration	increase N2 or H2	increase NH3
	decrease NH3	decrease N2 or H2

Volume- If increased, the reaction will shift toward the side with more gas so in this case LEFT for there are 4 moles of gas.

- If there is no gas, there is no effect.

Temperature- Increased, the reaction will proceed to the endothermic reaction(ΔH = positive)(Left)

- Decreased, the reaction will proceed to the exothermic direction (ΔH = negative)(Right)

Pressure- Increased, reaction will go toward the side with fewest molecules of gas(Right)

- Decreased, the reaction will go toward the side with more molecules of gas(Left)

Reaction Quotient(Q)

-By comparing K_c with Q you can determine the shift of the reaction
-K>Q will result in a forward shift
- K<Q will result in a backward shift

0.035 moles of SO_2, 0.500 moles of SO_2Cl_2, and 0.080 moles of Cl_2 are combined in an evacuated 5.00 L flask and heated to 100°C. What is Q before the reaction begins? Which direction will the reaction proceed in order to establish equilibrium?

$SO_2Cl_2(g) \leftrightarrows SO_2(g) + Cl_2(g)$ $K_c = 0.078$ at 100°C

$$Q_c = \frac{[SO_2][Cl_2]}{[SO_2Cl_2]}$$

0.500 mole SO_2Cl_2/5.00 L = 0.100 M SO_2Cl_2

0.035 mole SO_2/5.00 L = 0.070 M SO_2

0.080 mole Cl_2/5.00 L = 0.016 M Cl_2

$$Q_c = \frac{(0.070)(0.016)}{(0.100)} = 0.011$$

0.078 (K) > 0.011 (Q)
Therefore this reaction will shift right