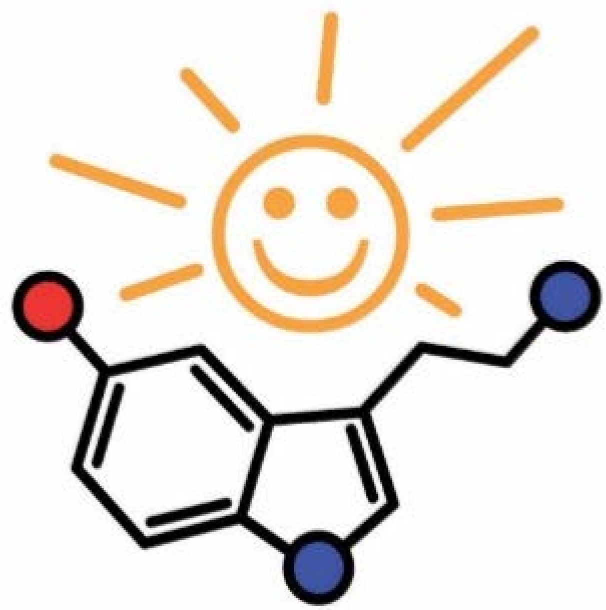
Student Name: ____

Class Period: ____

RRHS AP Chemistry



Unit 7:

Intro Equilibrium

Table of Contents

Class Info

Website: **magicalchemists.weebly.com/ap-chemistry.html**
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Remind Info:

A Day: Text **@APLestikA** to **81010**
B Day: Text **@APLestikB** to **81010**

Mastering Chemistry (Pearson textbook/homework)

https://www.pearsonmylabandmastering.com/northamerica/masteringchemistry/

You will need to get your course ID and access code from your teacher!

PERIODIC TABLE OF THE ELEMENTS

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.10	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.91	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 ***La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.21	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.02	89 **†Ac** 227.03	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (277)	109 **Mt** (268)	110 **Ds** (271)	111 **Rg** (272)							

***Lanthanide Series**	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.5	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
†Actinide Series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

2nd Marking Period: October-December 2022

Sunday	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday
9 Oct.	**10** Student Holiday/ Staff Development	**11** **A**	**12** **PSAT**	**13** **B**	**14** **C** End of 1st MP	15
16	**17** **A**	**18** **B**	**19** **A**	**20** **B**	**21** **A**	22
23	**24** Student Holiday/ Staff Development	**25** **B**	**26** **A**	**27** **B**	**28** **A**	29
30	**31** **B**	**1 Nov.** **A**	**2** **B**	**3** **A**	**4** **B**	5
6	**7** **A**	**8** **B**	**9** **A**	**10** **B**	**11** **A**	12
13	**14** **B**	**15** **A**	**16** **B**	**17** **A**	**18** **B**	19

Sunday	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday
20 Nov	**21** FALL BREAK	**22** FALL BREAK	**23** FALL BREAK	**24** FALL BREAK	**25** FALL BREAK	26
27	**28** **A**	**29** **B**	**30** **A**	**1 Dec.** **B**	**2** **A**	3
4	**5** **B**	**6** **A**	**7** **B**	**8** **A**	**9** **B**	10
11	**12** Exams 2, 1	**13** Exams 6, 5	**14** Exams 3, 4	**15** Exams 7, 8	**16** **BREAK BEGINS!**	17
18	**19** **WINTER**	**20** **BREAK**	**21** **WINTER**	**22** **BREAK**	**23** ☺	24
25	**26** ☺	**27** **WINTER**	**28** **BREAK**	**29** **WINTER**	**30** **BREAK**	31

AP® CHEMISTRY EQUATIONS AND CONSTANTS

Throughout the exam the following symbols have the definitions specified unless otherwise noted.

L, mL	=	liter(s), milliliter(s)	mm Hg	=	millimeters of mercury
g	=	gram(s)	J, kJ	=	joule(s), kilojoule(s)
nm	=	nanometer(s)	V	=	volt(s)
atm	=	atmosphere(s)	mol	=	mole(s)

ATOMIC STRUCTURE

$$E = h\nu$$
$$c = \lambda\nu$$

E = energy
ν = frequency
λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s
Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
Avogadro's number $= 6.022 \times 10^{23}$ mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}, \text{ where } a\,A + b\,B \rightleftarrows c\,C + d\,D$$

$$K_p = \frac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$K_b = \frac{[OH^-][HB^+]}{[B]}$$

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$
$$= K_a \times K_b$$

$$\text{pH} = -\log[H^+],\ \text{pOH} = -\log[OH^-]$$

$$14 = \text{pH} + \text{pOH}$$

$$\text{pH} = \text{p}K_a + \log\frac{[A^-]}{[HA]}$$

$$\text{p}K_a = -\log K_a,\ \text{p}K_b = -\log K_b$$

Equilibrium Constants

K_c (molar concentrations)
K_p (gas pressures)
K_a (weak acid)
K_b (weak base)
K_w (water)

KINETICS

$$\ln[A]_t - \ln[A]_0 = -kt$$

$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$$

$$t_{1/2} = \frac{0.693}{k}$$

k = rate constant
t = time
$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{\text{total}} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_A + P_B + P_C + \ldots$$

$$n = \frac{m}{\boldsymbol{M}}$$

$$\text{K} = {}^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

$$\text{Molarity, } M = \text{moles of solute per liter of solution}$$

$$A = abc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
$\boldsymbol{M}$ = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
a = molar absorptivity
b = path length
c = concentration

Gas constant, $R = 8.314 \text{ J mol}^{-1} \text{K}^{-1}$
$= 0.08206 \text{ L atm mol}^{-1} \text{K}^{-1}$
$= 62.36 \text{ L torr mol}^{-1} \text{K}^{-1}$
1 atm = 760 mm Hg = 760 torr
STP = 273.15 K and 1.0 atm
Ideal gas at STP = 22.4 L mol^{-1}

THERMODYNAMICS / ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$
$$= -RT \ln K$$
$$= -nFE^\circ$$

$$I = \frac{q}{t}$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)

Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

AP Chem: Effective Study Skills Tips and Tricks!

Study smarter, not harder. ☺

What to Do	What **NOT** to Do
Be **ACTIVE** in while learning/studying: • Close your booklet and try problems on your own with just a periodic table and formula chart! Only check your answer/work when you've finished, or you can't go any farther. • Use flashcards (physical or digital) • Struggle with challenging problems and keep trying, even if you're stuck initially (or convinced you're doing it wrong)	Be passive while learning/studying: • Re-read over your booklet and practice problems you've already completed • Ask your friend or look up the answer if you don't immediately know how to do the problem
Focus when studying • Decrease distractions while studying; don't read texts, check social media, or watch Netflix while studying. Put your phone out of sight/hearing.	Multitask • Study while checking/writing texts, checking social media, and/or watching Netflix. • Keep your computer or tv on in the background
Use **Intensity** when studying • You control the effort that you apply in your work! 30 minutes of high focus, high intensity study can be better than 2 hours of unfocused, low energy multi-tasking.	Low intensity/low effort • Look over problems and try them "in your head" but then just look up the answer • Use flashcards but don't try to recall the info on the other side before looking at the answer
Space out studying over time • Study a little bit of chemistry most days • Start long-term homework (like Mastering Chem or lab reports) the day they're assigned, and work a little bit every day or two • Less is more! Spaced practice studying is more effective than LONG hours of "studying" with multitasking and little focus.	Cram • Only study for quizzes/tests the night before • Start Mastering Chem or your lab report only 1-2 days before it's due • Study for many hours at a time all at once
Interleave your Studying • Study more than just one type of problem; mix it up and jump between different concepts • Review and practice old units while studying (especially important since AP Chem assessments are cumulative!)	One Concept Studying • Study only one type of problem, and practice those problems over and over • Don't review older content or units while studying

Test Yourself!
• The best way to prepare for a test is to take a test!
  o Time yourself while trying practice problems
  o Access only the AP Periodic Table and Formula Chart when practicing problems
	Open Notes Practice
	• Use your notes, friends, and/or the internet while trying practice problems
	• Give yourself unlimited time for each problem

AP Chemistry Big Ideas

BIG IDEA 6 - Any bond or intermolecular attraction that can be formed can be broken. These processes are in a dynamic competition, sensitive to initial conditions and external perturbations.

- Enduring Understanding 6.A: Chemical equilibrium is a dynamic, reversible state in which rates of opposing processes are equal.
- Enduring Understanding 6.B: Systems at equilibrium are responsive to external perturbations, with the response leading to a change in the composition of the system.
- Enduring Understanding 6.C: Chemical equilibrium plays an important role in acid-base chemistry and solubility.
- Enduring Understanding 6.D: The equilibrium constant is related to temperature and the difference in Gibbs free energy between reactants and products.

Equilibrium: Let's Get Balanced!

Dynamic Equilibrium: A Quick Review

1. The concentrations of all reactants and products remain ________________ with time.
2. The reaction is proceeding in the forward **and** reverse direction simultaneously and at the ___________ rate.
3. All macroscopic variables (such as concentration, partial pressure, and temperature) do ______ change over time, so nothing _______________ to be happening.
4. When equilibrium has been reached, ____ = *K*.
5. **Equilibrium is temperature dependent**! Change the temperature, change _________ of products to reactants.

The Equilibrium Constant, *K* (the Law of Mass Action): relates the concentrations of reactants and products at equilibrium at a given temperature.

For the general reaction: aA + bB ⇋ cC + dD

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

- Each concentration is raised to the power of its stoichiometric coefficient in the balanced equation.
- **Note:** Only use ____________ and ___________________ substances (**NO solids or pure liquids!)**
- There are NO ______________ for the equilibrium constant, *K* (they cancel out).

Let's Practice!

1. The decomposition of aqueous $Ca(HCO_3)_2$ is allowed to come to equilbrium at 298 K. The measured equilibrium concentrations are $[Ca(HCO_3)_2]$ = 0.025 M and $[CO_2]$ = 2.78 M. What is the value of the equilibrium constant at this temperature?

   $$Ca(HCO_3)_2(aq) \rightleftharpoons CaCO_3(s) + H_2O(l) + CO_2(g)$$

2. Consider the following reaction:

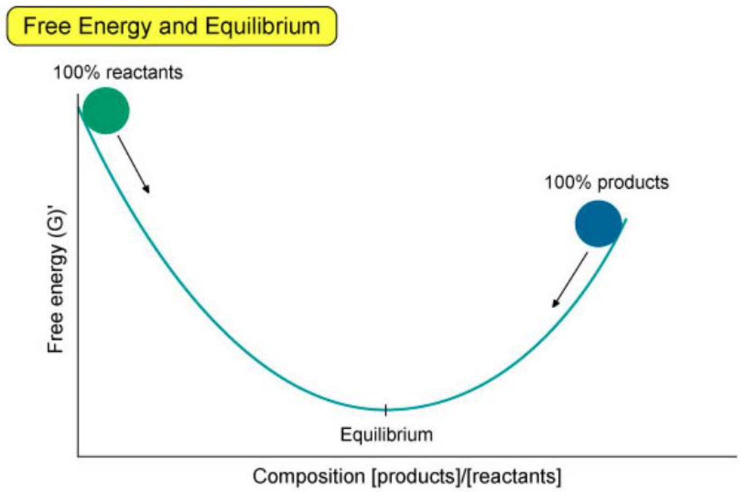
   $$2\ COF_2(g) \rightleftharpoons CO_2(g) + CF_4(g) \qquad K_c = 2.00 \text{ at } 1000^\circ C$$

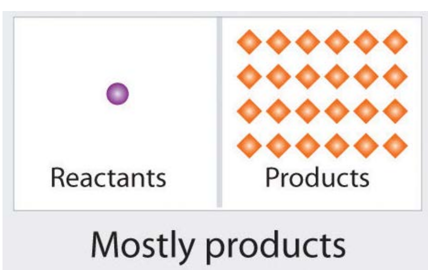
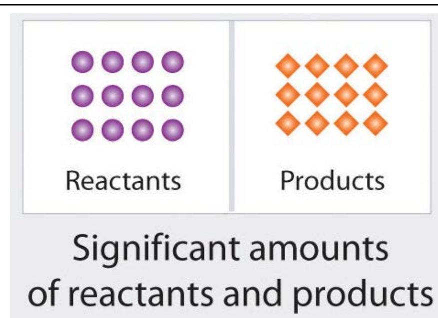
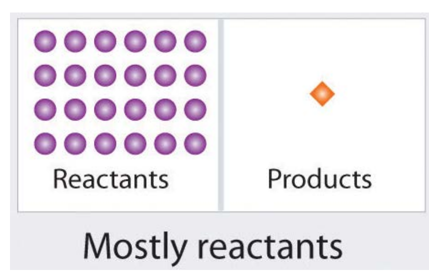
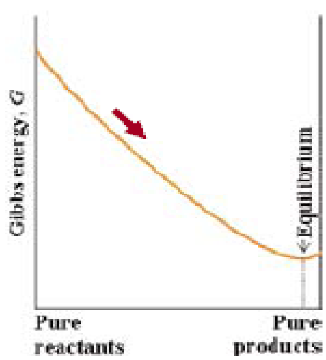
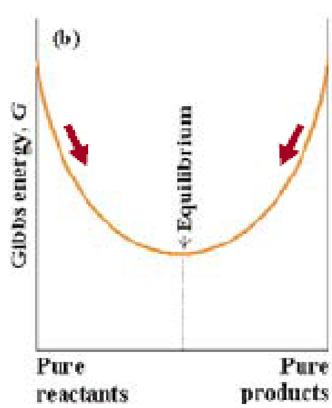
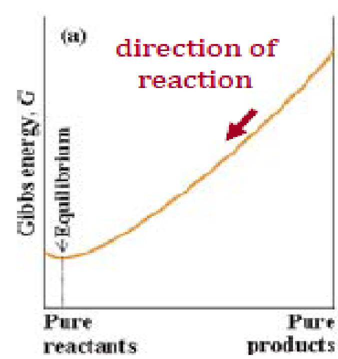
   In an equilibrium mixture, the concentration of COF_2 is 0.255 M and the concentration of CF_4 is 0.118 M. What is the equilibrium concentration of CO_2?

What does *K* mean?

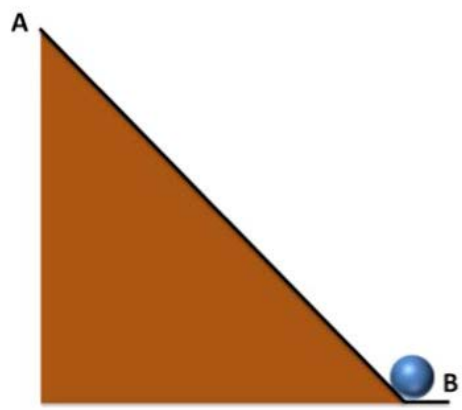
The equilibrium constant, K, tells you:

- the ratio of products to reactants when a given reaction reaches its ___________ free energy state and "stops".



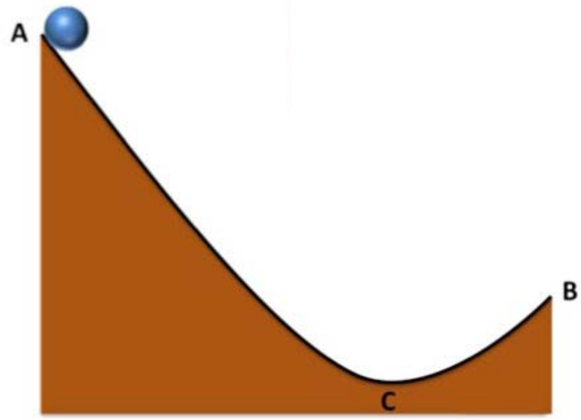
Large *K* $K >> 1$	**Intermediate *K*** $K \approx 1$	**Small *K*** $K << 1$
Product-favored	Neither	Reactant-Favored
		
Forward reaction is thermodynamically favorable 		Reverse reaction is thermodynamically favorable 

Completion Reactions vs Equilibrium Reactions



Complete Reaction: Complete consumption of reactants.

$$A \rightarrow B$$



Equilibrium: Reaction achieves equilibrium by reaching the lowest state of energy.

$$A \rightleftharpoons B$$

Completion reactions:	Equilibrium reactions:
• Non-reversible • ________ of reactants convert to products • Example: combustion	• Reversible • Reaction will occur until lowest energy state is reached • Example: weak acid dissolution

You can think of K as measuring how close a reaction will go towards 100% completion:

- High K = reaction came pretty close to completely turning reactants into products
- Low K = reaction did NOT come close to turning all reactants into products, mostly reactants just hung around

Kinetic Control: when a reaction is thermodynamically favorable ($K > 1$), but products form sooooooo slowly that it appears the reaction isn't happening (i.e., kinetically unfavorable)

Example Question: For a given reaction $A \rightleftharpoons B$, $\Delta H = -2.3$ kJ/mol and $\Delta S = +25.9$ J/mol K. A student tries the reaction in lab, letting the reaction run over time. After two days. she measures the composition of her sample and discovers that it is primarily composed of reactants. Which of the following best explains why this might occur?

a. This reaction is only thermodynamically favorable at low temperatures, and the student's lab must be too warm for the reaction to occur.

b. This reaction has a small equilibrium constant, *K*, and thus the [reactants] will be much greater than [products] when the reaction system reaches its most stable state.

c. This reaction has an extremely high activation energy, and thus the reaction rate is so slow no products will be observed despite its thermodynamic favorability.

d. This reaction is only thermodynamically favorable at high temperatures, therefore the lab conditions are too cold for the reaction to occur in measurably quantities.

Pesky Subscripts: Different designations of K

In order to convey additional information, the equilibrium constant *K* may also have a _________________ to give information about the type of reaction being studied.

However, in _________ cases K is still defined as the ratio of products to reactants!

Type of Reaction		Reaction	Equilibrium Expression, *K*
General	K_{eq}	aA + bB ⇋ cC + dD	$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$
Concentration	K_c		
Pressure (only for gases!)	K_p	aA(g) + bB(g) ⇋ cC(g) + dD(g)	$K_p = \frac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$
Dissolving a solid in water	K_{sp}	AB(s) ⇋ A^+(aq) + B^-(aq)	$K_{sp} = [A^+][B^-]$

K_p = Ratio of Partial Pressures

For the general reaction: aA + bB ⇌ cC + dD

$$K_p = \frac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b} = \frac{(P_{products})^{coefficient}}{(P_{reactants})^{coefficient}}$$ where P is the partial pressure of the gas

*Note: K_c can be converted into K_p, but that conversion is no longer AP tested! ;D

Let's Practice!

1. Given the following reaction, with K_p = 109 at 25ºC: 2 NO(g) + Br_2(g) ⇌ 2 NOBr(g)

   If the equilibrium partial pressure of bromine gas is 0.0159 atm and the equilibrium partial pressure of NOBr is 0.0768, calculate the equilibrium partial pressure of NO:

2. The reaction for the formation of nitrosyl chloride: 2 NO(g) + Cl_2(g) ⇌ 2 NOCl(g) was studied at 25°C. The pressures at equilibrium were found to be P_{NOCl} = 1.2 atm, P_{NO} = 0.050 atm and P_{Cl2} = 0.30 atm. Write the equilibrium expression, K_p, for this reaction and calculate its value at 25°C.

Manipulating Reactions and the Effect on K

1. Stoichiometric Coefficients: If you ____________________ the coefficients in the equation by a factor, __________ the equilibrium constant to the same factor to get the equilibrium constant for the reaction.

$$2\,\mathrm{X} + \mathrm{Y} \leftrightarrow 2\,\mathrm{Z} \qquad K$$

$$\mathrm{X} + \frac{1}{2}\mathrm{Y} \leftrightarrow \mathrm{Z} \qquad K^{\frac{1}{2}}$$

$$4\,\mathrm{X} + 2\,\mathrm{Y} \leftrightarrow 4\,\mathrm{Z} \qquad K^2$$

2. Reversing Equations: When the equation is written in __________________, take the ______________________ of K to get the equilibrium constant for the reaction.

$$2\,\mathrm{X} + \mathrm{Y} \leftrightarrow 2\,\mathrm{Z} \qquad K$$

$$2\,\mathrm{Z} \leftrightarrow 2\,\mathrm{X} + \mathrm{Y} \qquad \frac{1}{K}$$

3. Adding Equations: If you ________ two or more chemical equations to get the overall reaction (like in Hess's Law), __________________ the respective K's to get the equilibrium constant for the reaction.

$$K_{total} = K_1 \times K_2 \times K_3 \ldots$$

Let's Practice!

1. The Haber Process is a famous industrial method for producing ammonia from nitrogen and hydrogen gases:

$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g) \qquad K_c = 3.8 \times 10^4 \text{ at } 127°C$$

   a. Calculate the value of the equilibrium constant, K_c, at 127°C for the reaction: $2\,NH_3(g) \rightleftharpoons N_2(g) + 3\,H_2(g)$

   b. Calculate the value of K_c at 127°C for this reaction: $\frac{1}{2}\,N_2(g) + 3/2\,H_2(g) \rightleftharpoons NH_3(g)$

   c. Calculate the value of K_c at 127°C for this reaction: $6\,NH_3(g) \rightleftharpoons 3\,N_2(g) + 9\,H_2(g)$

THE REACTION QUOTIENT, Q: When you need to know the answer to the question, "Is the system at equilibrium?"

A: The answer can be ________ or _______!

For the general reaction: aA + bB ⇌ cC + dD

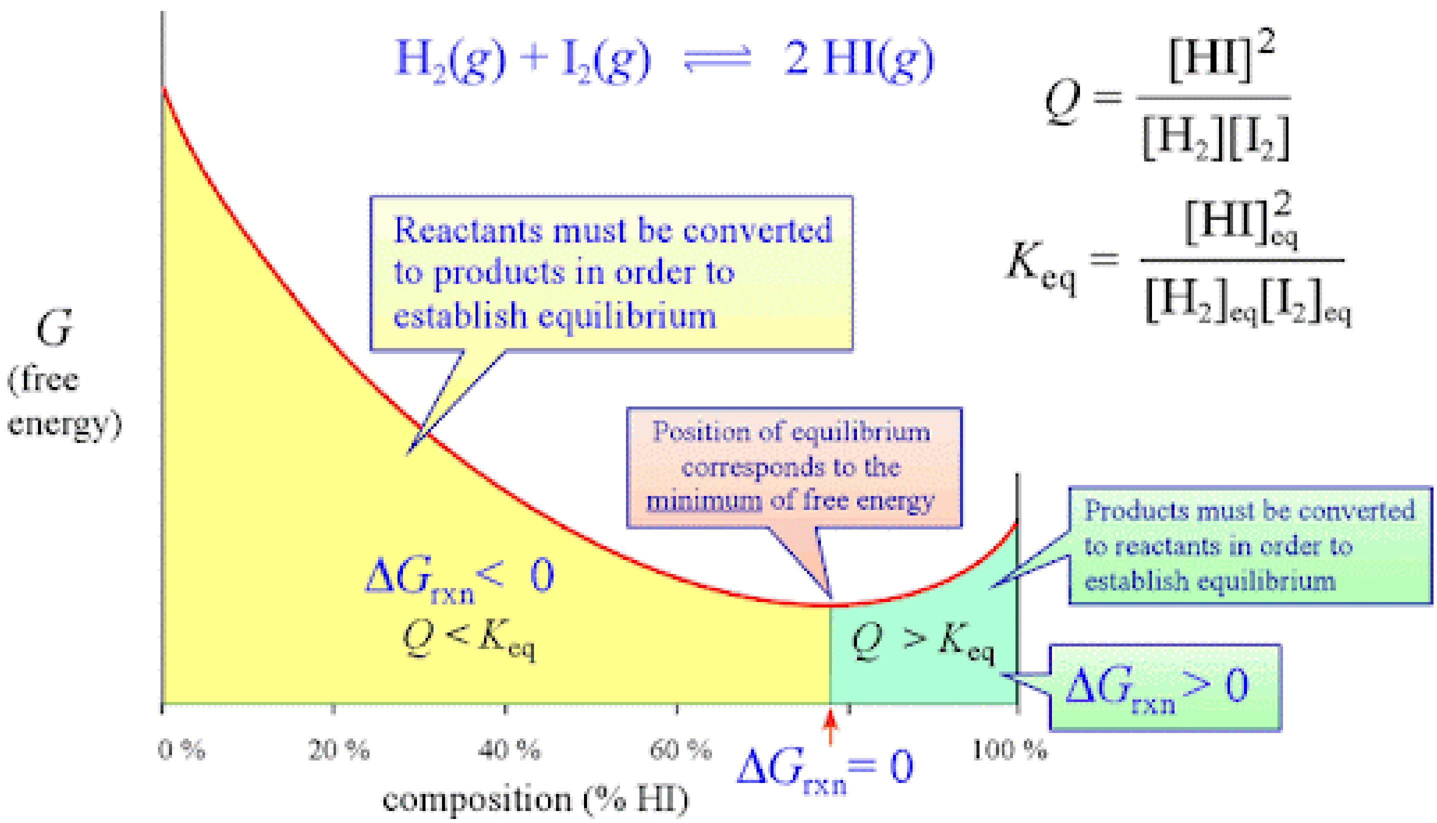
$$Q = \frac{[C]^c[D]^d}{[A]^a[B]^b} \quad \text{or} \quad \frac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$$

<u>Reminder</u>:

- Q has the appearance of K (same exact ratio!)
- But… Q can be calculated at _____ point in the reaction progress, not only at equilibrium!

What does *Q* mean?

1. If ***K*** ____ ***Q***, system not at equilibrium: forward reaction is favored (shift right) to make Q = K.
2. If ***K*** ____ ***Q***, the system is at equilibrium.
3. If ***K*** ____ ***Q***, system not at equilibrium: reverse reaction is favored (shift left) to make Q = K.



The Kinetics of Equlbrium

1. If equilibrium is approached from the left (starting with ______________),
    a. the rate of the <u>forward</u> reaction ____ to a constant, non-zero rate (i.e. it slows down over time until equilibrium is reached).
    b. the rate of the <u>reverse</u> reaction ____ to a constant, non-zero rate (i.e. it speeds up over time until equilibrium is reached).
2. If equilibrium is approached from the right (starting with ______________),
    a. the rate of the <u>forward</u> reaction ____ to a constant, non-zero rate (i.e. it speeds up over time until equilibrium is reached).
    b. the rate of the <u>reverse</u> reaction ____ to a constant, non-zero rate (i.e. it slows down over time until equilibrium is reached).
3. Time required to reach equilibrium does _______ depend on the equilibrium constant, *K*!
4. Regardless of initial conditions, at a given _________________ a reaction will reach equilibrium with the same ratio of products to reactants.

In Summary

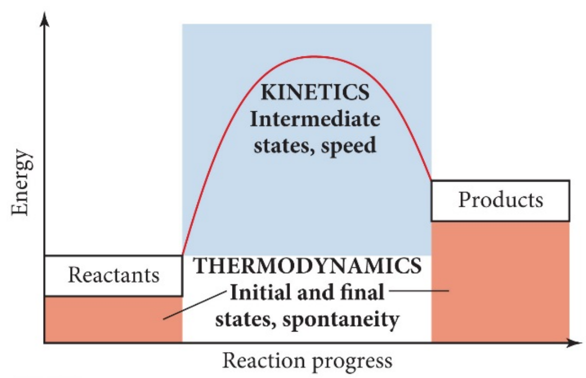
Current conditions	$K > Q$	$K \approx Q$	$K < Q$
change needed for system to reach equilibrium	shift right (make more products)	already at equilibrium	shift left (make more reactants
reaction rates	forward > reverse reaction rate (until equilibrium reached)	forward ≈ reverse reaction rate	forward < reverse reaction rate (until equilibrium reached)

Thermodynamics:

- Will a reaction happen spontaneously?
- Determined by ΔG (i.e., combo of ΔH and ΔS)

Kinetics:

- How fast will a reaction happen?
- Determined by activation energy, E_a, and temperature



Equilibrium vs Kinetics: k vs *K*!

Equilibrium	**Kinetics**
K = equilibrium constant	*k* = ___________ constant

What we **<u>can</u>** determine about *k* using *K*:

- Relative rates of forward and reverse reactions (by comparing K vs Q)

What we **<u>can't</u>** determine about *k* using *K*:

- Absolute rates of forward and reverse reactions

You CANNOT compare the rate of one reaction to another by comparing their *K* values!

Let's Practice!

1. At 1000 K, the value of K_p for the reaction $2\ SO_3(g) \rightleftharpoons 2\ SO_2(g) + O_2(g)$ is 0.338. Predict which direction the reaction will proceed toward equilibrium if the initial partial pressures are: P_{SO3} = 0.16 atm, P_{SO2} = 0.41 atm, and P_{O2} = 2.5 atm.

2. For the synthesis of ammonia at 500°C, $N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g)$, the equilibrium constant is 6.0×10^{-2}. Predict the direction in which direction the system will shift to reach equilibrium if $[NH_3]_{\text{initial}} = 1.0 \times 10^{-4}$ M, $[N_2]_{\text{initial}} = 5.0$ M, and $[H_2]_{\text{initial}} = 1.0 \times 10^{-2}$ M at 500°C.

3. The equilibrium partial pressure of Br_2 is 4.00 atm and that of NOBr is 8.00 atm.

$$2\ NO(g) + Br_2(g) \leftrightarrow 2\ NOBr(g) \qquad K_p = 100$$

Using the equation above, determine the equilibrium partial pressure of nitrogen monoxide, NO, at equilibrium.

a. 0.400 atm b. 0.566 atm c. 1.77 atm d. 5.83 atm

4. The value of the equilibrium constant, K_c, at 25°C is 8.1 for the following reaction: $2\ SO_3(g) \rightleftharpoons 2\ SO_2(g) + O_2(g)$. What must happen for the reaction to reach equilibrium if the initial concentrations of all three species was 2.0 M?

   a. The rate of the forward reaction would increase, and $[SO_3]$ would decrease.
   b. The rate of the reverse reaction would increase, and $[SO_2]$ would decrease.
   c. Both the rate of the forward and reverse reactions would increase, and the value for the equilibrium constant would also increase.
   d. No change would occur in either the rate of reaction or the concentrations of any of the species.

5. For the reaction $2\ X(g) + Y(g) \leftrightarrow 2\ Z(g)$, $K_c = 4.0 \times 10^4$. Determine the value of the equilibrium constant, K_c, for the following reaction:

$$2\ Z(g) \leftrightarrow 2\ X(g) + Y(g) \qquad K_c = ??$$

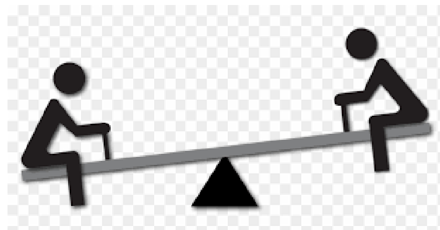
a. 2.5×10^{-5} b. 2.5×10^{-4} c. 4.0×10^{-5} d. 4.0×10^{-4}

Le Châtelier's Principle

If a "stress" (_____________) is applied to a system at equilibrium,

equilibrium will shift in the direction that will partially relieve that stress.

Nothing can ever completely reverse the effects of stress! (Tell me about it, am I right? ;P)



Changes in Concentration

- Increasing the concentration of a reactant causes system to shift right to make more ______________________.
- Increasing the concentration of a product causes system to shift left to form more ________________________.
- Decreasing the concentration of a product causes system to shift right to form more ______________________.
- Decreasing the concentration of a reactant causes system to shift left to make more ______________________.

Important note: adding or subtracting the amount of a **pure solid or **liquid** will NOT cause a shift.**

A change in concentration of reactants or products will _______ affect the value of *K*.

$$\mathrm{A}(aq) + \mathrm{B}(aq) \rightleftarrows \mathrm{C}(aq) + \mathrm{D}(aq)$$

Stress	Effect	Stress	Effect
Add A or B	reaction will shift ______________ to make more _________________	Add C or D	reaction will shift ______________ to make more _________________
Remove A or B	reaction will shift ______________ to make more _________________	Remove C or D	reaction will shift ______________ to make more _________________

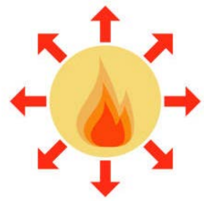
Example: $\mathrm{N_2}(g) + 3\ \mathrm{H_2}(g) \rightleftarrows 2\ \mathrm{NH_3}(g)$

What will happen to the equilibrium system above when the following stresses are applied?

Stress	Effect	Stress	Effect
Add N_2	reaction will shift ______________ to make more _________________	Add H_2	reaction will shift ______________ to make more _________________
Add NH_3	reaction will shift ______________ to make more _________________	Remove NH_3	reaction will shift ______________ to make more _________________

Changes in Temperature

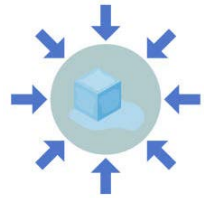
Changes in temperature may easily be treated as changes in concentration if you think of heat as a reactant (endothermic reaction) or product (exothermic reaction).



Exothermic Reactions: Heat is a ______________

$$A + B \rightleftharpoons C + 56\text{ kJ}$$

- Increasing temperature causes system to shift left. *K* will _____________________.
- Decreasing temperature causes system to shift right. *K* will ____________________.



Endothermic Reactions: Heat is a ______________

$$D + E + 43\text{ kJ} \rightleftharpoons F + G$$

- Increasing temperature causes system to shift right. *K* will ___________________.
- Decreasing temperature causes system to shift left. *K* will ___________________.

****Only a change in _______________ will change the value of *K*.****

Let's Practice!

1. Given the reaction: $2\,SO_3(g) \rightleftharpoons O_2(g) + 2\,SO_2(g) \quad + 32.5\text{ kJ}$

   What will happen to the equilibrium system above when the following stresses are applied?

Stress	Effect on Reaction System	Effect on K
Decrease the temperature	reaction will shift ______________ to make more _________________	The value of K will:
Add SO_3	reaction will shift ______________ to make more _________________	The value of K will:
Increase the temperature	reaction will shift ______________ to make more _________________	The value of K will:

2. When a certain amount of nitrogen gas and hydrogen gas are placed in a 1.5 L evacuated container and heated to 773 K, ammonia gas, NH_3, is formed according to the following equation: $N_2(g) + H_2(g) \rightleftharpoons 2\,NH_3(g)$. When equilibrium is established, 0.565 mol of $NH_3(g)$ is present in the flask. When the same reaction is carried out at 298 K, the number of moles of $NH_3(g)$ is much larger than 0.565 mol. Is the forward reaction endothermic or exothermic? Justify your answer.

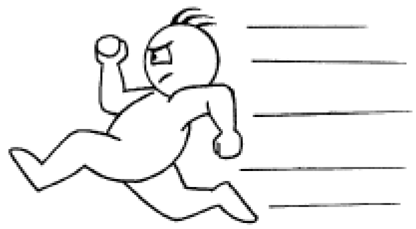
Oh Shift: Le Châtelier and Rate of Reaction

A shift to the left or right does **NOT** say anything about the ___________ of the reaction!

For example, consider the reaction: A + B ⇌ C + heat

If the temperature of this system was increased, equilibrium would shift to the left. This does _______ mean that the rate will be slower! It simply means that a new equilibrium will be reached which has more A and B and less C.

Increasing temperature causes equilibrium to be reached more ________________ (regardless of shift!)



Changes in Pressure

Changes in pressure only affect equilibrium systems that have ________________ reactants and/or products.

- Increasing the pressure of a gaseous system causes system to shift to the side with ___________ gas particles.
- Decreasing the pressure of a gaseous system causes system to shift to the side with ___________ gas particles.
- If system has the ___________ number of moles of gas on both sides, changes in pressure have _____ effect.
- Adding an inert gas does _______ affect the location of equilibrium, since the partial pressures of the gases in the reaction are not affected.

Changes in Volume: The opposite of changes in pressure.

- Increasing the volume of a gaseous system causes system to shift to the side with ___________ gas particles.
- Decreasing the volume of a gaseous system causes system to shift to the side with ___________ gas particles.

A change in pressure or volume will _______ affect the value of *K*.

Addition of a Catalyst: No effect on equilibrium position!

- Adding a catalyst increases the rate of both the forward and reverse reaction ____________.
- Equilibrium will be reached more quickly. ;D

Example: $\mathrm{N_2}(g) + 3\ \mathrm{H_2}(g) \rightleftarrows 2\ \mathrm{NH_3}(g)$

What will happen to the equilibrium system above when the following stresses are applied?

Stress	Effect	Stress	Effect
Increase pressure	reaction will shift ______________ to make more _________________	Ar(g) added	reaction will shift ______________ to make more _________________
Increase volume	reaction will shift ______________ to make more _________________	Decrease pressure	reaction will shift ______________ to make more _________________

Let's Practice!

Consider the following equilibrium equation:

$$Ca(HCO_3)_2(aq) \rightleftharpoons CaCO_3(s) + H_2O(l) + CO_2(g) \qquad \Delta H_{rxn} = +2.1 \text{ kJ/mol}$$

A vessel contains $Ca(HCO_3)_2$, $CaCO_3(s)$, $NO_2(g)$, $H_2O(l)$, and $CO_2(g)$ at equilibrium. Predict how each of the following stresses will affect the variables specified.

Stress	**Shift?**	**Effect on K?**	**Other Effect**
CO_2 is added.			The concentration of $Ca(HCO_3)_2$ will:
$CaCO_3$ is removed.			The concentration of CO_2 will:
The volume is halved.			The amount of $CaCO_3$ will:
He(g) is added.			The amount of $Ca(HCO_3)_2$ will:
The temperature is decreased.			The concentration of $CaCO_3$ will:
The pressure is decreased.			The concentration of $CaCO_3$ will:
$CaCO_3$ is added.			The concentration of $Ca(HCO_3)_2$ will:
A catalyst is added.			The amount of CO_2 will:
The volume is increased.			The concentration of $Ca(HCO_3)_2$ will:
The temperature is increased.			The amount of CO_2 will:

A Visual Summary of Le Châtelier's Principle

Le Châtelier: How to FRQ

1. Identify stress (change) to the system.
2. Identify effect of stress on the equilibrium system: shift right, shift left, or no shift?
3. Explain how the shift you identified will counteract the original stress to re-establish equilibrium
4. Connect your explanation to the question asked (aka, answer the question. ;D)

Example FR Explanations

Example #1: If the temperature of an exothermic reaction changes from 25^{o}C to 50^{o}C, what is the effect on equilibrium constant, K?

- *Stress?* Increasing temperature
- *Effect of stress?* Shift left (or shift towards reactants)
- *Counteract stress?* Use up added heat
- *Answer question (effect on K):* increasing [reactants], decreasing [products] decreases value of K.

Final Answer: Increasing the temperature of an exothermic reaction will cause the system to shift left to use up the added heat and re-establish equilibrium. This shift increases the concentration of reactants and decreases the concentration of products, therefore the value of K decreases.

Example #2: For the reaction A(g) + B(g) $\rightleftharpoons$ C(g), adding more A(g) to the reaction vessel will have what effect on the concentration of B(g)?

- *Stress?* Adding A(g) = increasing [A], a reactant
- *Effect of stress?* Shift right (or shift towards products)
- *Counteract stress?* Use up some of the added A(g)
- *Answer question (effect on concentration of B):* decrease [B], because using up some A to create products also uses up B.

Final Answer: Adding more A(g) will increase the concentration of A, a reactant, and so the system will shift right to use up some of the extra A by making more product. Using up some A to make products will also use up some B, and so the concentration of B will decrease as the system shifts to re-establish equilibrium.

Example #3: For the reaction A(g) + B(s) $\rightleftharpoons$ C(g), removing some B(s) from the reaction container will have what effect on the reaction mixture?

- *Stress?* No stress occurs: solids are not included in the equilibrium constant expression, K
- *Effect of stress?* n/a
- *Counteract stress?* n/a
- *Answer question (effect on rxn mixture):* no effect on location of equilibrium for the reaction mixture

Final Answer: Removing some B(s) will have no effect, because solids are not included in the equilibrium constant expression, *K*, and thus don't affect the location of equilibrium for the reaction mixture.

Free Response Practice!

1. In the following reaction at chemical equilibrium, what is the effect on the reaction mixture if the volume of the reaction container is decreased? Explain.

$$N_2(g) + O_2(g) \rightleftharpoons 2\ NO(g) \qquad \Delta H° = 181\ kJ$$

2. Consider the following endothermic reaction at chemical equilibrium. When the temperature of the reaction mixture is increased, what is the effect on $CaCO_3(s)$? Explain.

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

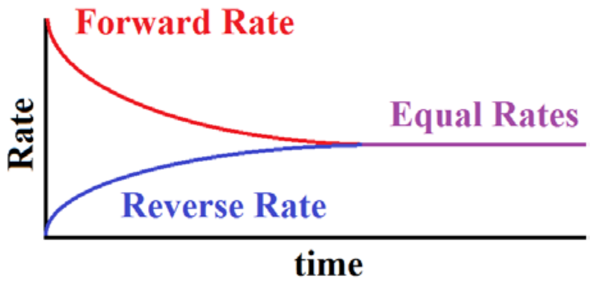
3. Adding additional KCl will have what effect on the reaction mixture, if the mixture was originally at chemical equilibrium? Explain.

$$2\ KClO_3(s) \rightleftharpoons 2\ KCl(s) + 3\ O_2(g)$$

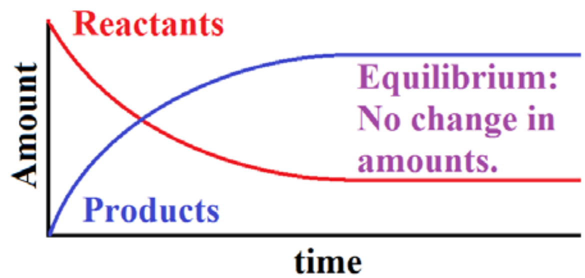
Quick Review: Equilibrium Graphs

Equilibrium systems can be represented by two different graphs:

Rate of Reaction vs Time



[Reactants], [Products] vs Time



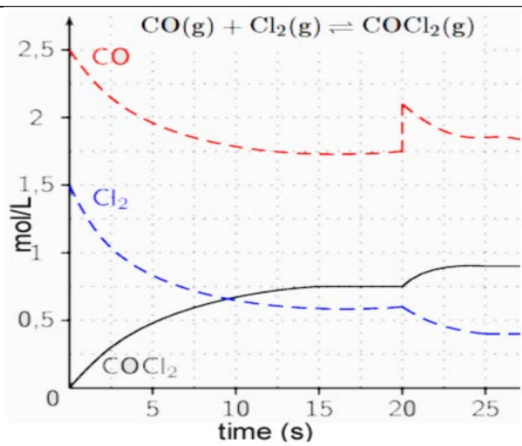
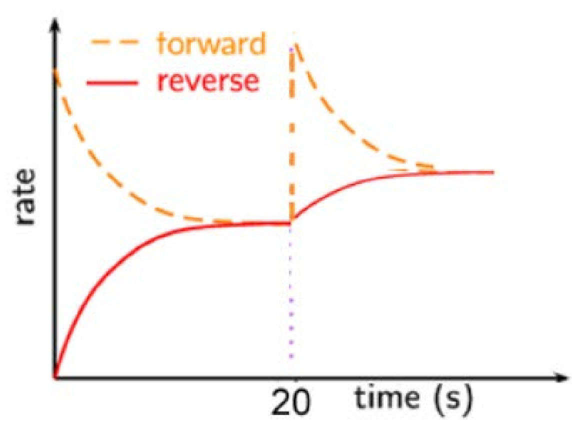
How can you use each graph to identify when equilibrium is reached?

a. **Rate of Reaction vs Time:** both forward and reverse reaction rates are ____________________.

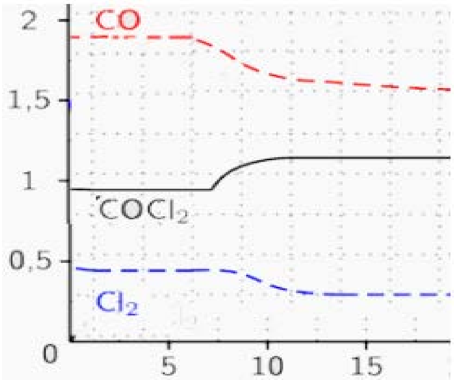
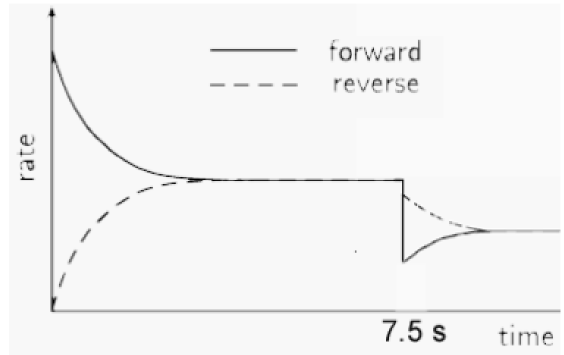
b. **[Reactants], [Products] vs Time:** concentrations of reactants and products are _____________________, but not necessarily equal.

Le Châtelier's Principle: Interpreting Equilibrium Graphs

1) Applied stress: Change in Concentration

Concentration vs. Time	**Rate** vs. Time
 	
Visual clues: 1) ___________ in the concentration of **one** reactant or product a. Spike UP if concentration _______________ b. Spike DOWN if concentration ____________ 2) ___________ change in the concentration of all other reactants and/or products ➔ until all concentrations are constant again	Visual clues: 1) ___________ in the rate of **one** direction a. Spike UP if concentration ______________ b. Spike DOWN if concentration ___________ 2) ___________ change in the other rate 3) until both rates are equal again

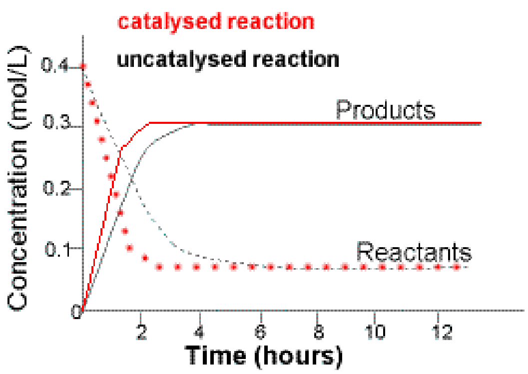
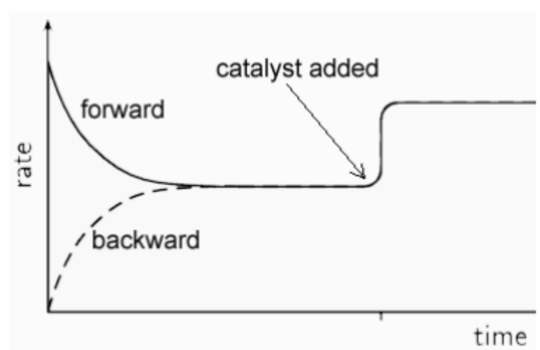
2) **Applied Stress:** Change in Temperature

Concentration vs. Time	**Rate** vs. Time
	
Visual clues: 1) __________ change in the concentration of **all** reactants and products ➔ until all concentrations are <u>constant</u> again 2) Direction of shift determines which side increases/decreases concentration	Visual clues: 1) __________ in the rate of **one** direction a. Spike UP if temperature ___________ b. Spike DOWN if temperature ___________ 2) __________ change in the other rate 3) until both rates are <u>equal</u> again

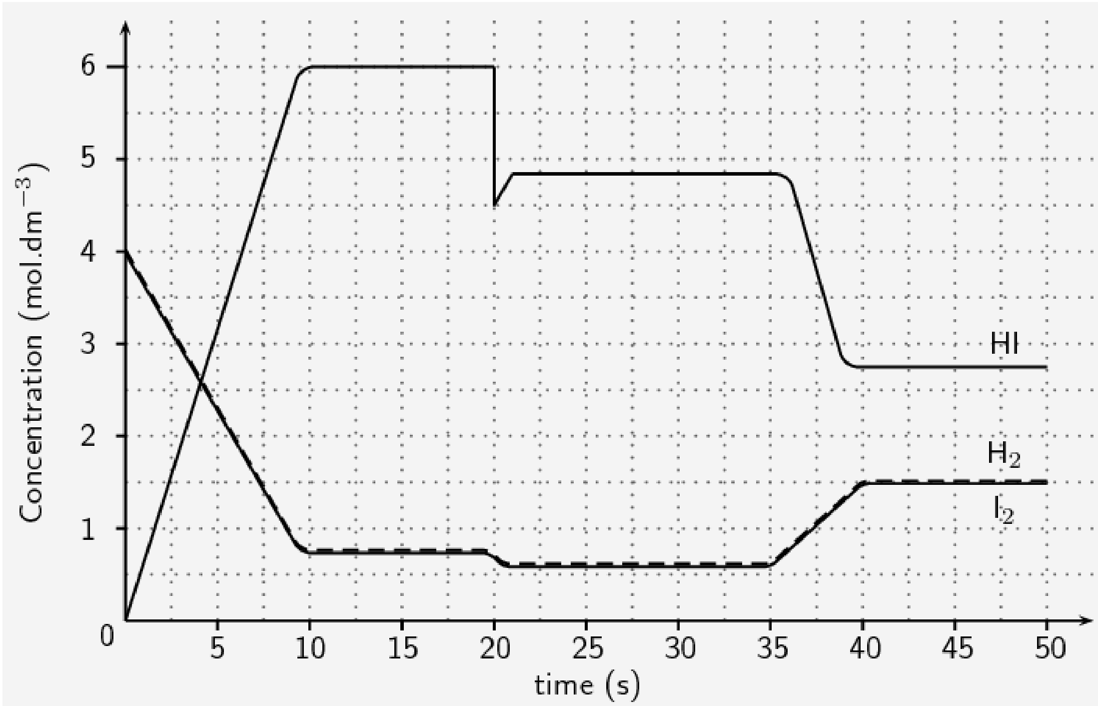
Reminder: **Nothing can ever completely undo the stress on the system!**

Shifts predicted by Le Châtelier's principle only _________________ counteract the stress.

3) **Applied stress:** Addition of a Catalyst

Concentration vs. Time	**Rate** vs. Time
	
Visual clues: ➔ System reaches equilibrium at the _________ concentrations but an ____________ time	Visual clues: ➔ Rate of forward and backward reactions increases ____________________

Graphing Practice!



Use the chemical equilibrium below and the graph above to answer the questions that follow.

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

4. How many seconds does it take for the system to reach equilibrium the first time? _______________

5. Calculate the value of the equilibrium constant at the time identified in #4.

6. Identify a stress to the system that would cause the effects shown at t = 20 s. Explain.

7. If the change at t = 35 s is due to an increase in temperature, is the reaction exothermic or endothermic? Justify your answer.

Multiple Choice Practice!

Use the following information to answer questions 1−3.

The following reaction is found at equilibrium at 25°C: $2\ SO_3(g) \rightleftharpoons O_2(g) + 2\ SO_2(g)$ ΔH = −198 kJ/mol

1. Which of the following would cause the reverse reaction to speed up?

   a. Adding more SO_3
   b. Lowering the temperature
   c. Raising the pressure
   d. Removing some SO_2

2. Which of the following would cause a reduction in the value for the equilibrium constant?

   a. Increasing the amount of SO_3
   b. Raising the temperature
   c. Reducing the amount of O_2
   d. Lowering the temperature

3. If initially only SO_3 was added to the reaction vessel, what is true about the following values as the system approached equilibrium?

   a. $\Delta G > 0$ and $Q > K$ b. $\Delta G < 0$ and $Q > K$ c. $\Delta G > 0$ and $Q < K$ d. $\Delta G < 0$ and $Q < K$

Use the following information to answer questions 4−5.

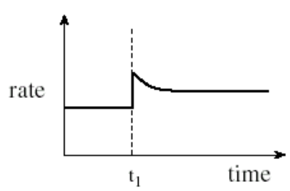
$$CO_2(g) + H_2(g) \rightleftharpoons H_2O(g) + CO(g)$$

4. Which two stresses will each cause the equilibrium to shift to the left?

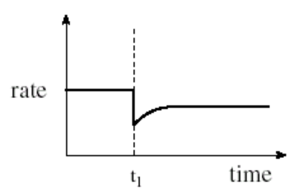
   a. increase $[H_2]$, increase [CO]
   b. increase $[CO_2]$, decrease [CO]
   c. decrease $[H_2]$, increase $[H_2O]$
   d. decrease $[CO_2]$, decrease $[H_2O]$

5. Which of the following graphs represents the forward rate of reaction when $H_2O_{(g)}$ is added to the above equilibrium at time t = 1 ?

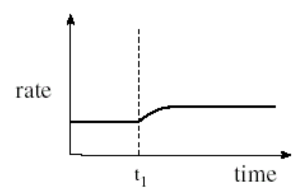
A.



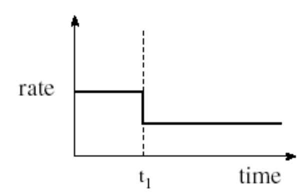
B.



C.



D.



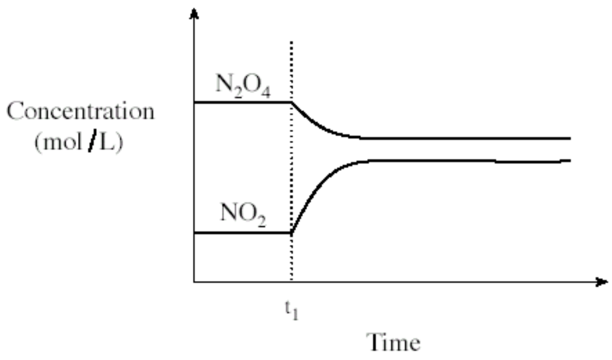
6. Consider the following: $2\ SO_3(g) \rightleftharpoons O_2(g) + 2\ SO_2(g)$. Initially, SO_3 is added to an empty flask. How do the rate of the forward reaction and $[SO_3]$ change as the reaction proceeds to equilibrium?

	Forward Rate	$[SO_3]$		Forward Rate	$[SO_3]$
a.	decreases	increases	c.	increases	increases
b.	decreases	decreases	d.	increases	decreases

7. The following reaction is found at equilibrium: $Ni(s) + 4\ CO(g) \rightleftharpoons Ni(CO)_4(l)$ ΔH = −160.8 kJ/mol
   Which of the following will cause this equilibrium to shift to the left?

   a. add some CO
   b. remove some $Ni(CO)_4$
   c. decrease the volume
   d. increase the temperature

8. Consider the following reaction at chemical equilibrium: $N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$. At time t_1. heat is applied to the system. Which of the following best describes the equilibrium reaction and the change in K_c?

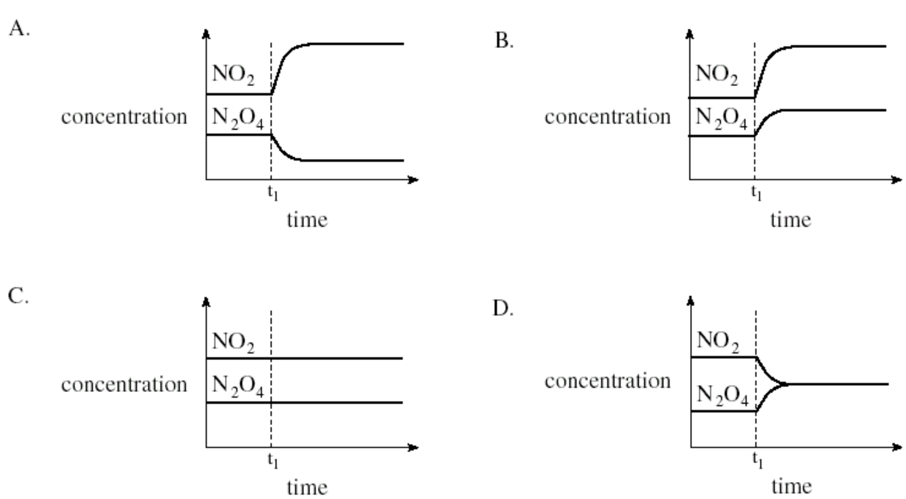


a. exothermic and K_c increases
b. exothermic and K_c decreases
c. endothermic and K_c increases
d. endothermic and K_c decreases

9. A galvanic cell is constructed based on the following reaction: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$. The observed voltage was found to be 0.95 volt instead of the standard cell potential, E°, of 1.10 volts. Which of the following could correctly account for this observation?

A. The cell had been running for a period of time.
B. The standard free energy of the cell, ΔG°, is negative.
C. The Cu^{2+} solution was more concentrated than the Zn^{2+} solution.
D. The Zn(s) anode had been reduced in mass.

10. Consider the following equilibrium: $N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$. Which of the following shows the relationship between concentration and time as a result of adding a catalyst at time t = 1?



11. The following reaction is found at equilibrium: $N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g)$. When the volume of the system is decreased, the equilibrium shifts:

a. left since the reverse rate is greater than the forward rate.
b. left since the forward rate is greater than the reverse rate.
c. right since the reverse rate is greater than the forward rate.
d. right since the forward rate is greater than the reverse rate.

RICE Tables: Delicious, Delicious Equilibrium

Many equilibrium calculations are solved using ___________ tables: this is an organization method used to clearly track what happens as a reaction established equilibrium.

Tasty RICE

- **R** = Balanced ____________
- **I** = _____________ concentrations (or pressures) for each species in the reaction mixture
- **C** = _____________ in concentrations (or pressures) for each species as the system moves towards equilibrium: everything changes stoichiometrically!
- **E** = ______________ concentrations (or pressures) of each species when the system reaches equilibrium

Note: If the amounts are given in moles BE WARY − you must convert to _________________ (M).

Example: Consider the reaction $2\ SO_3(g) \rightleftharpoons O_2(g) + 2\ SO_2(g)$

A reaction container has an initial $[SO_3]$ of 0.020 M. At equilibrium, $[SO_2]$ = 0.012 M. Calculate the value of the equilibrium constant.

Reaction	$2\ SO_3(g)$ $\rightleftharpoons$	$O_2(g)$ +	$2\ SO_2(g)$
Initial concentration			
Change			
Equilibrium concentration			

Now you try! Consider the following reaction: $2\ CH_4(g) \rightleftharpoons C_2H_2(g) + 3\ H_2(g)$

A reaction mixture at 1700 °C initially contains $[CH_4]$ = 0.115 M. At equilibrium, the mixture contains $[C_2H_2]$ = 0.035 M. What is the value of the equilibrium constant?

Questionable Q: Predicting the Direction of a Reaction with Yummy RICE!

If you're given initial concentrations of _______ reactants and products, you need to calculate _____ to predict the direction of shift ____________ using a RICE table to calculate equilibrium concentration.

Example: Consider the reaction: $2\ NO(g) \rightleftharpoons N_2(g) + O_2(g)$ $K_c = 1.00 \times 10^{-6}$ At 298 K, 2.25 moles of NO, 0.0749 moles of N_2, and 0.0750 moles of O_2 are placed into a 1.50 L flask and allowed to reach equilibrium. Calculate the equilibrium concentrations for each substance listed below at 298 K.

First, calculate Q. Which direction will the reaction shift?

Now you can make a RICE table! (Knowing which direction reaction shifts allows you to determine correct signs for x.)

Reaction			
Initial concentration			
Change			
Equilibrium concentration			

Now you can write your equilibrium expression and plug in with the values from above.

Finally, what are your equilibrium concentrations?

Check yourself!

Now you try! Consider the reaction: $I_2(g) + Cl_2(g) \rightleftharpoons 2\ ICl(g)$, $K_p = 81.9$ (at 25°C).

A reaction mixture at 25°C initially contains P_{I_2} = 0.100 atm, P_{Cl_2} = 0.100 atm, and P_{ICl} = 0.100 atm. Find the equilibrium partial pressures of I_2, Cl_2, and ICl at this temperature.

Multiple Choice Practice!

1. Ammonia and oxygen react according to the following equilibrium: $4\ NH_3(g) + 3\ O_2(g) \rightleftharpoons 2\ N_2(g) + 6\ H_2O(g)$. A 2.0 liter flask is initially filled with 8.0 mol of oxygen and 6.0 mol of ammonia, and 6.0 mol of H_2O are present when the system reaches equilibrium. How much oxygen is present at equilibrium?

   a. 2.0 mol O_2 b. 3.0 mol O_2 c. 5.0 mol O_2 d. 6.0 mol O_2

RICE Approximations: Simplifying the Math

Sometimes RICE math can get complicated. Let's try an example!

1. What are the equilibrium concentrations of each chemical species in a 0.15 M solution of HCN at 22°C?

$$HCN(aq) \rightleftharpoons H^+(aq) + CN^-(aq) \qquad K_c = 4.9 \times 10^{-10} \text{ at } 22^\circ C$$

But wait! You should never actually have to use the ________________ equation on the AP Chemistry test. (Mastering Chem, however, is a different story… ;D)

But how can we solve a problem of this type without doing piles of math?

The "X is Negligible" Approximation

There are two common problem types where this approximation is valid:

a. Very ___________ K values (where $K < 10^{-5}$ or Initial Molarity ≥ 1000 x K) when starting from reactants: in this case, barely any products will be made.

b. Very ___________ K values (where $K > 10^{5}$) when starting from products: in this case, barely any reactants will be made.

Note: The assumption about x being negligible only works when x is less than ______ of the initial concentration, so you can check the validity of your assumption after solving for x!

Let's try that first problem again, this time using the "x is negligible" approximation.

1. What are the equilibrium concentrations of each chemical species in a 0.15 M solution of HCN at 22°C?

$$HCN(aq) \rightleftharpoons H^+(aq) + CN^-(aq) \qquad K_c = 4.9 \times 10^{-10} \text{ at } 22^\circ C$$

2. For the reaction $N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$, which of the following conditions allows you to conclude that the change in concentration, x, is negligible? (Multiple correct answers!)

   a. At a constant temperature, $[NO_2] = 1.0$ M and $K_c = 4.0 \times 10^{-7}$.

   b. At a constant temperature, $[N_2O_4] = 1.0$ M and $K_c = 4.0 \times 10^{-7}$.

   c. At a constant temperature, $[N_2O_4] = 1.0$ M and $K_c = 4.0 \times 10^{6}$.

   d. At a constant temperature, $[NO_2] = 1.0$ M and $K_c = 4.0 \times 10^{6}$.

3. Consider the equilibrium reaction: $I_2(g) \rightleftharpoons 2\ I(g)$, where $K_c = 5.6 \times 10^{-12}$. If 1.58 moles of of $I_2(g)$ are placed into a 3.5 L reaction vessel, what are the equilibrium concentrations of the reactant and the product?

What if x isn't negligible??

The Perfect Square: if keeping "−x" is necessary, see if the problem is a perfect _____________ and thus easy to solve using delicious algebra.

Example: Consider the reaction: $H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g)$, $K_p = 64$ (at 25°C). A reaction mixture at 25°C initially contains $P_{H_2} = P_{I_2} = 0.20\ \text{atm}$. Find the equilibrium partial pressures of H_2, I_2, and HI.

Now you try!

1. If the K_p for the following reaction is 2.4 x 10^{-9} and the initial concentration of CO_2 is 2.00 atm, what are the partial pressures of the substances at equilibrium?

$$C(s) + CO_2(g) \rightleftharpoons 2CO(g)$$

2. A gas, XY(g), decomposes according to the following reaction: $2\,XY(g) \rightleftharpoons X_2(g) + Y_2(g)$, $K_p = 230$. A sample of each of the gases is places in a previously evacuated container, and the initial partial pressures of the gases are shown in the table below:

Gas	Initial Partial Pressure (atm)
XY	0.010
X_2	0.20
Y_2	2.0

If the temperature of the reaction mixture is held constant, in which direction will the reaction proceed? Explain.

3. For the reaction $N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$, 1.0 mol of $N_2O_4(g)$ is placed in a 10.0 L reaction vessel at a constant temperature. $K_c = 4.0 \times 10^{-7}$ for this temperature. Find the equilibrium concentrations.

Multiple Choice Practice!

4. Here is a general reaction with a K value of 16: $A(aq) + B(aq) \rightleftharpoons 2\ C(aq)$.
   Initially, [A] = [B] = 2.0 M. Solve for the equilibrium concentration of each substance.
   a. [A] = [B] = 0.67 M, [C] = 1.3 M
   b. [A] = [B] = 1.6 M, [C] = 0.88 M
   c. [A] = [B] = 0.67 M, [C] = 2.7 M
   d. [A] = [B] = 0.50 M, [C] = 3.0 M

5. Consider the following equilibrium: $Cl_2(g) + 2\ NO(g) \rightleftharpoons 2\ NOCl(g)$. Initially, the reaction was started by adding 10.0 *M* NOCl gas to a reaction vessel. At equilibrium, [NO] = 2.0 *M*. What is the value of K_c?
   a. 0.63
   b. 2.3
   c. 4.0
   d. 16

6. The reaction below came to equilibrium at a temperature of 100ºC. At equilibrium the partial pressure due to NOBr was 4 atm, the partial pressure due to NO was 4 atm, and the partial pressure due to Br_2 was 2 atm. What is the equilibrium constant, K_p, for this reaction at 100ºC?

$$2\ NOBr(g) \rightleftharpoons 2\ NO(g) + Br_2(g)$$

a. ¼ b. ½ c. 1 d. 2

7. Consider the following: $2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g)$. Initially, 0.030 mol SO_2 and 0.030 mol O_2 are placed into a 1.0 L container. At equilibrium, there is 0.020 mol O_2 present. What is the $[SO_2]$ at equilibrium?

a. 0.010 mol/L b. 0.020 mol/L c. 0.030 mol/L d. 0.040 mol/L

Use the information below to answer #8−10.

Reaction 1:	$NOBr(g) \rightleftharpoons NO(g) + ½\ Br_2(g)$	$K_c = 3.4 \times 10^{-2}$
Reaction 2:	$2\ NOCl(g) \rightleftharpoons 2\ NO(g) + Cl_2(g)$	$K_c = 1.6 \times 10^{-5}$
Reaction 3:	$2\ NO(g) + 2\ H_2(g) \rightleftharpoons N_2(g) + 2\ H_2O(g)$	$K_c = 4.0 \times 10^{6}$
Reaction 4:	$N_2(g) + O_2(g) \rightleftharpoons 2\ NO(g)$	$K_c = 4.2 \times 10^{2}$

8. For a reaction involving nitrogen monoxide inside a sealed flask, the value for the reaction quotient (Q) was found to be 1.1×10^2 at a given point. If, after this point, the amount of NO gas in the flask increased, which reaction is most likely taking place in the flask?

a. reaction 1 b. reaction 2 c. reaction 3 d. reaction 4

9. Which of these reactions proceeds at the slowest rate?

a. reaction 1 b. reaction 2 c. reaction 3 d. cannot be determined

10. For reaction #3, equimolar amounts of N_2 gas and H_2O gas are allowed to come to equilibrium in a sealed reaction vessel. Which of the following must be true at equilibrium?

I. $[N_2]$ must be less than $[H_2O]$.
II. $[N_2]$ must be greater than $[H_2O]$.
III. $[NO]$ must be greater than $[H_2]$.

a. I only b. II only c. I and III d. II and III

Kitchen Sink Equilibria: Everything but the…

ΔG and K: Oh rats! $\Delta G^{o} = -RT \ln K$

R = universal gas constant = 8.314 J mol^{-1} K^{-1}

T = temperature in Kelvin

K = equilibrium constant

This equation can be rewritten to give: $K = e^{-\Delta G^{\circ}/RT}$

And of course, don't forget the connection between electrochemistry and equilibrium!

ΔG and E^{o}: I've had enounf! $\Delta G^{o} = -nFE^{o}$

n = number of moles of electrons

F = Faraday's constant = 96,485 coulombs per mole of electrons

1. Calculate Gibb's free energy, in kJ/mol_{rxn}, for the reaction of glucose and ATP to give glucose-6-phosphate and ADP, given that the equilibrium constant, *K*, is 5000. at a temperature of 38^{o}C.

2. Assume that iron and magnesium half-cells are suitable connected at 298 K and standard conditions and that both aqueous solutions are 1.00 M concentrations.

$Fe^{2+}(aq) + 2e^{-} \rightarrow Fe(s)$ $E^{o} = -0.45$ V

$Mg^{2+}(aq) + 2e^{-} \rightarrow Mg(s)$ $E^{o} = -2.37$ V

Calculate the cell potential, the free energy change (in kJ/mol), and the equilibrium constant for this voltaic cell.

Upping the Pressure

$$PV = nRT$$

P = pressure (in atm, torr, or mmHg)

R = universal gas constant (0.08206 L atm mol^{-1} K^{-1} or 62.36 L torr mol^{-1} K^{-1})

T = temperature in Kelvin

The ideal gas law can be used to calculate the partial pressure of a gas at equilibrium if you know the equilibrium concentration of the gas.

3. If Consider the following reaction:

   $$2\ COF_2(g) \rightleftharpoons CO_2(g) + CH_4(g) \qquad K_c = 2.00 \text{ at } 1000^{\circ}C$$

   In an equilibrium mixture, the concentration of COF_2 is 0.255 M and the concentration of CH_4 is 0.118 M. What is the partial pressure of CO_2 at equilibrium?

Equilibrium and Thermodynamic Favorability

4. For the gaseous equilibrium represented below, it is observed that greater amounts of PCl_3 and Cl_2 are produced as the temperature is increased.

   $$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$$

   a. What is the sign of ΔS° for the reaction? Justify.

   b. What change, if any, will occur in ΔG° for the reaction as the temperature is increased? Explain your reasoning in terms of Le Chatelier’s principle.

Multiple Choice Practice!

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

5. An electrolytic cell based on the reaction represented above was constructed from iron and copper half-cells. The observed voltage was found to be 0.59 volt instead of the standard cell potential, E°, of 0.78 volts. Which of the following could correctly account for this observation?

   A. The copper electrode was larger than the iron electrode.
   B. The solutions in the half-cells had different volumes.
   C. The Cu^{2+} solution was more concentrated than the Fe^{2+} solution.
   D. The Fe^{2+} solution was more concentrated than the Cu^{2+} solution.

$$H_2O(s) \rightarrow H_2O(l)$$

6. When ice is placed into warm water at room temperature, which of the following is true for the phase change shown above?

   a. $Q > K$
   b. ΔG is positive
   c. ΔH is negative
   d. ΔS is positive

$$2\ Al(s) + 3\ Mn^{2+}(aq) \rightarrow 2\ Al^{3+}(aq) + 3\ Mn(s)$$

7. A thermodynamically favorable cell, utilizing the reaction shown above, ran for 45 minutes. What happens to the measured voltage and why?

   A. The measured voltage decreases over time because deviations in concentration that bring the cell closer to equilibrium will decrease the magnitude of the cell potential.
   B. The measured voltage increases over time because deviations in concentration that bring the cell closer to equilibrium will increase the magnitude of the cell potential.
   C. The measured voltage increases over time because $[Mn^{2+}]$ increases as the cell runs.
   D. The measured voltage remains constant because E^{o}_{cell} is an intensive property.

8. Which of the following statements is true about the reaction below?

$$2\ Ag(s) + Cu^{2+}(aq) \rightleftharpoons 2\ Ag^{+}(aq) + Cu(s)\ K_{eq} = 2.7 \times 10^{-16}$$

   a. E° and ΔG° are both positive.
   b. E° and ΔG° are both negative.
   c. E° is positive and ΔG° is negative.
   d. E° is negative and ΔG° is positive.

9. Calculate the standard free energy of the following reaction at 25°C.

$$Fe^{2+}(aq) + Mg(s) \rightarrow Fe(s) + Mg^{2+}(aq) \qquad E^{o}_{cell} = 1.92\ V$$

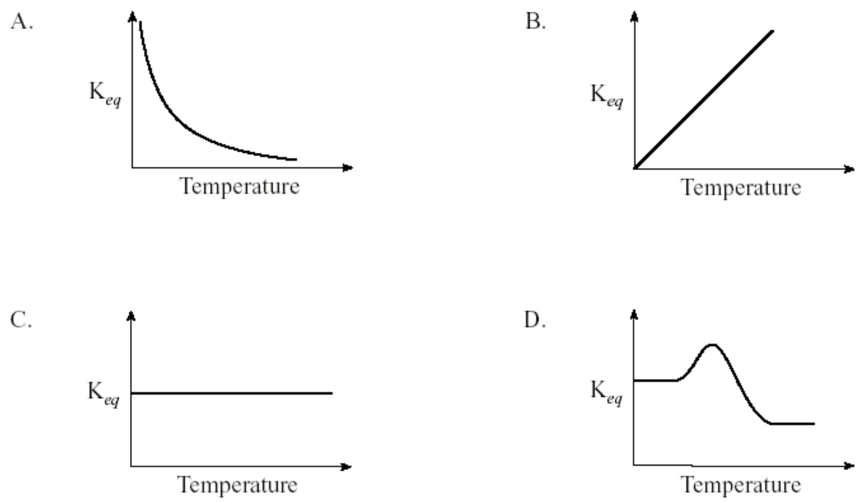
   a. 3.7×10^{5}
   b. 1.6×10^{3}
   c. -3.7×10^{5}
   d. -1.6×10^{3}

10. Which of the following would cause an increase in the potential of the voltaic cell described by the reaction below?

$$2\ Ag^{+}(aq) + Fe(s) \rightarrow 2\ Ag(s) + Fe^{2+}(aq)$$

a. Increasing $[Fe^{2+}]$
b. Adding more Fe(s)
c. Decreasing $[Fe^{2+}]$
d. Removing some Fe(s)

11. The relationship between K_{eq} and temperature for an exothermic reaction is represented by



12. Consider the following equilibrium:

$$2\ CO(g) + O_2(g) \rightleftarrows 2\ CO_2(g) \qquad K_c = 4.0 \times 10^{-10}$$

What is the value of K_c for $2\ CO_2(g) \rightleftarrows 2CO(g) + O_2(g)$?

a. 4.0×10^{10}
b. 2.0×10^{-5}
c. 2.0×10^{5}
d. 2.5×10^{9}

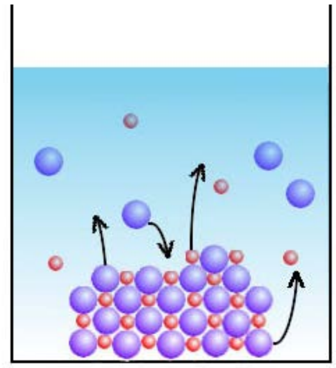
$$2\ NO(g) + Cl_2(g) \rightarrow 2\ NOCl(g) \qquad \Delta H = -77\ kJ/mol_{rxn}$$

13. Which of the following statements accurately describes the above reaction?

a. The entropy of the products exceeds that of the reactants.
b. NO(g) will always be the limiting reagent.
c. K will be greater than 1 at all temperatures.
d. The temperature of the surroundings will increase as this reaction progresses.

Solubility Equilibria: Dissolve All Your Troubles Away!

Up until now, we've pretended that compounds fall into one of two categories: 100% soluble or 100% insoluble. Actually, an equilibrium can exist between a partially soluble substance and its solution.



Solubility Product Constant (________)

Equilibrium expression for dissolving a solid into ions

Like all K values, K_{sp} is constant at a constant temperature.

$$\mathrm{AgCl}(s) \rightleftarrows \mathrm{Ag}^{+}(aq) + \mathrm{Cl}^{-}(aq)$$

When the rate at which a solid dissolves into ions is ____________________ to the rate at which ions precipitate back to solid, the system has reached equilibrium.

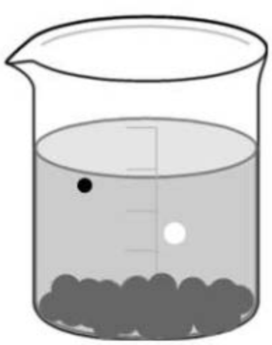
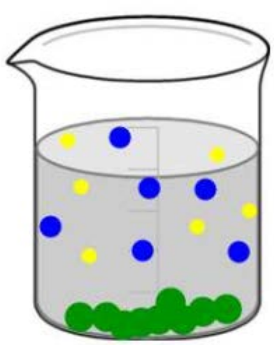
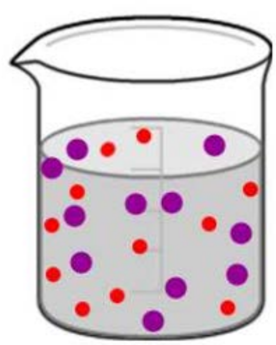
Let's Practice! Write the *Ksp* expression for each of the following dissolutions:

Salt	Dissociation reaction	K_{sp} Expression
K_2CO_3		
Al_2S_3		

Wait a sec, K_{sp} only has ________________________! But why?

<u>Reminder</u>: There are three types of 100% soluble ions that you have to memorize:

<u>Always soluble</u>: __________ metal cations, ______, ______

Small K_{sp} (K << 1)	Larger K_{sp}	100% Soluble
• Only a small amount of solid dissociates into ions • Lower solubility	• More solid dissociates into ions • Higher solubility	• All solid dissociates into ions • Dissolves to completion
		

K_{sp} Calculations

ALWAYS write the balanced dissociation reaction and K_{sp} expression before solving!

Example: When silver sulfide dissolves in pure water, $[S^{2-}] = 3.4 \times 10^{-17}$ mol/L at 25°C. Calculate its K_{sp} value.

Now you try!

1. What is the K_{sp} value of silver phosphate at 25°C, if $[PO_4^{3-}] = 1.8 \times 10^{-18}$ M when dissolved in pure water?

2. What is the concentration of chloride ions that will form in a solution of $PbCl_2$ in pure water? The K_{sp} of lead (II) chloride at 25°C is 1.17×10^{-5}.

But wait! The "____" term for dissociation always refers to the amount of _______________ that will dissolve (since the stoichiometric coefficient of the salt will always be _____ in the dissociation reaction. This term has a special name!

<u>Solubility</u> "S" (aka <u>Molar Solubility</u>) = "x" in your K_{sp} RICE Table

How much of a __________ will dissolve per _____ L of solution (Units: M = mol/L)

Solubility is an equilibrium position and therefore ______ change (for example, if you change the number of ions in solution, this will shift the equilibrium position and thus, the solubility).

- Larger molar solubility values suggest _________ dissociation into ions and <u>greater</u> solubility.
- Smaller molar solubility values suggest _________ dissociation into ions and <u>lower</u> solubility.

<u>Important Note about molar solubility</u>: When comparing solubility of given compounds, compare molar solubility values, ______ K_{sp} values

- *Exception*: can directly compare molar solubility values if and only if the compounds compared dissociate into the ___________ number of ions

Now we can answer problems using the same math we used on the previous page. The terminology will be different, but the calculations are the same! Let's try some:

3. The molar solubility of barium fluoride is at 25°C is 2.45×10^{-5} M. Calculate K_{sp}.

4. Calculate the molar solubility of nickel (II) carbonate, which has a K_{sp} of 1.4×10^{-7} at 25°C.

How Much Will Dissolve?

You can use molar solubility to determine how many ____________ of a solid will dissolve in a quantity of water!

5. How many grams of iron (II) hydroxide can dissolve in 500. mL of water? Its $K_{sp} = 4.87 \times 10^{-17}$ at 25°C.

6. The molar solubility of $PbCl_2$ in pure water is 1.43×10^{-2} M at 25ºC.
   a. Write the equilibrium constant expression for the dissolving of $PbCl_2(s)$.
   b. How many grams of $PbCl_2$ can dissolve into 200. mL of pure water?
   c. What change could be made to increase amount of $PbCl_2$ that will dissolve?

Multiple Choice Practice!

7. The solubility product, K_{sp}, of AgCl is 1.8×10^{-10}. Which of the following expressions is equal to the molar solubility of AgCl?

   a. $(1.8 \times 10^{-10})^2$ molar
   b. $\frac{1.8 \times 10^{-10}}{2}$ molar
   c. 1.8×10^{-10} molar
   d. $\sqrt{1.8 \times 10^{-10}}$ molar

8. Which of the following is equal to the solubility product, K_{sp}, of Ag_2CO_3?

   a. $K_{sp} = [Ag^+][CO_3{}^{2-}]$
   b. $K_{sp} = [Ag^+][CO_3{}^{2-}]^2$
   c. $K_{sp} = [Ag^+]^2[CO_3{}^{2-}]$
   d. $K_{sp} = [Ag^+]^2[CO_3{}^{2-}]^2$

9. If the solubility of BaF_2 is equal to x, which of the following expressions is equal to the solubility product, K_{sp}, of BaF_2?

   a. x^2
   b. $2x^2$
   c. $2x^3$
   d. $4x^3$

10. In a saturated solution of Na_3PO_4, $[Na^+] = 0.30$ M. What is the molar solubility of Na_3PO_4?

    a. 0.10 M
    b. 0.30 M
    c. 0.60 M
    d. 0.90 M

11. What is the maximum mass of AgCl can dissolve in 100. mL of pure water at 25°C? The molar solubility of AgCl is 1.3×10^{-5}.

    a. 1.9×10^{-4} g
    b. 1.9×10^{-5} g
    c. 9.1×10^{-9} g
    d. 9.1×10^{-6} g

Solubility Lab

Although all compounds have a characteristic solubility in water at a given temperature, some families of compounds are more soluble than others and it is useful to understand the general rules of solubility. If the solubility of a substance is greater than 0.1 mol/L, we call the substance soluble. If the solubility is less than 0.01 mol/L, we call the substance insoluble, and if the solubility is between 0.01 mol/L and 0.1 mol/L, we call the substance slightly soluble or sparingly soluble. Solubility equilibrium is typically studied in saturated solutions of slightly soluble compounds.

Part 1: Precipitation of Slightly Soluble Compounds

Reaction #1:

1. Place one drop of 0.10 M lead (II) nitrate, $Pb(NO_3)_2$, in a clean well.
2. Add one drop of 0.10 M sodium bromide, NaBr, to the same well. Mix well by tapping/ flicking the well plate.

Observations:
Complete reaction:
Net-ionic reaction:
Spectator ions:

Reaction #2:

1. Place one drop of 0.10 M lead (II) nitrate, $Pb(NO_3)_2$, in a clean well.
2. Add one drop of 0.10 M sodium iodide, NaI, to the same well. Mix well by tapping/ flicking the well plate.

Observations:
Complete reaction:
Net-ionic reaction:
Spectator Ions:

Part 2: Relative Solubility

Challenge: Determine which of the two lead-containing precipitates formed in Part 1 is least soluble.

Materials: a clean well plate, 0.10 M $Pb(NO_3)_2$ solution, 0.10 M NaBr solution, and 0.10 M NaI solution.

Experimental Design: In the space below,

1. Briefly outline the experimental procedure you will follow (include all equipment and solutions).
2. Explain how the data collected will allow you to determine which of the two lead precipitates is least soluble.
3. *Check the answers to #1 and #2 with your teacher. Once they sign off, you may start your experiment!*

Observations/Data:

Results: Which precipitate is least soluble? Explain how your data supports your conclusion.

Part 3: Relative Solubility with Numbers

Use the information provided in the table below to answer the following questions.

	K_{sp}	Molar solubility (mol/L)
$PbBr_2$	6.6×10^{-6}	
PbI_2		1.3×10^{-3}

1. Complete the table above by filling in the missing values. Show all math in the space below.

2. Using the values in the completed table, determine which of the two precipitates is least soluble. Explain.

3. Does your answer to #2 support the results of your experiment in Part 2? Explain.

Part 4: One-pot Solubility

Consider the information in the table below when answering the following questions.

	Precipitate color	K_{sp}
AgCl	milky white	1.8×10^{-10}
Ag_2CO_3	white with yellow tint	8.1×10^{-12}

1. The following ions are all present in solution at the same time: $Ag^+(aq)$, $Cl^-(aq)$, and $CO_3^{2-}(aq)$. Using the information provided in the table above and your knowledge and solubility equilibria, predict which of the two precipitates above will form under these conditions. Justify your prediction.

2. Design an experiment to test the prediction you made in question #1. Your available materials are one well plate, 0.10 M $AgNO_3(aq)$, 0.10 M $Na_2CO_3(aq)$, and 0.10 M $NaCl(aq)$. Record your observations/data below.

Observations/Data:

Results: Which precipitate formed when ions of Ag^+, Cl^-, and CO_3^{2-} were present? Explain. Did your data support your prediction in question #1?

Clean-Up: Empty the solutions into the designated waste heavy metals container. Wash the well plates thoroughly with soap and water, rinse with tap water, then with distilled water. Clean up all other materials **and wash your hands** with soap and water.

The Common Ion Effect

<u>Reminder</u>: solubility can change if you change reaction conditions!

- Le Châtelier's principle predicts that a salt will become __________ soluble in a solution that already contains one of its own ions already dissolved: what's known as a _______________ ion.
- The presence of a common ion acts like increasing the concentration of a _________________ ion in the salt dissolution, causing the system to shift ___________ to establish equilibrium (towards the ________ side).

Example:

1. Circle any of the following compounds that contain a common ion to $MgCl_2$:

   AgF_2 $NaCl$ $Mg(OH)_2$ $AlCl_3$ Al_2S_3

2. Which of the compounds above, if present in the solution in equal concentration, would reduce the solubility of $MgCl_2$:
   a. the most? Why?

   b. the least? Why?

The Effect of pH on Solubility

The common ion effect predicts that when a salt contains ions that can act as an acid or a base, the solubility of that salt will be affected by changes in ________.

Example: $$Fe(OH)_3 \rightleftharpoons Fe^{3+}(aq) + 3\ OH^{-}(aq)$$

3. Will iron (III) hydroxide be more, less, or equally soluble in a **<u>basic</u>** solution (when compared to its solubility in pure water)? Explain.

4. Will iron (III) hydroxide be more, less, or equally soluble in an **<u>acidic</u>** solution (when compared to its solubility in pure water)? Explain.

Common Ion Calculations

5. Calculate the molar solubility of nickel (II) carbonate in a solution containing 0.100 M $NaCO_3$. The K_{sp} of $NiCO_3$ is 1.4×10^{-7} at 25°C.

6. Calculate the molar solubility of SrF_2 in a solution containing 0.400 M NaF. The K_{sp} of SrF_2 is 7.9×10^{-10} at 25°C.

Common Ion Applications: Achieving a Desired Concentration

You can use the common ion effect to control the ____________________ of the **non-common** ion in solution! Let's see how this chemical magic works.

7. In pure water at 25°C, the K_{sp} of $CaCO_3$ is 3.8×10^{-9}. $Ca(NO_3)_2$ is added to 1.00 L of a saturated solution of $CaCO_3$ at 25°C until the $[CO_3^{2-}]$ is reduced to 2.3×10^{-7} M. How many moles of $Ca(NO_3)_2$ are dissolved in solution at the point when $[CO_3^{2-}] = 2.3 \times 10^{-7}$ M? (Assume the added $Ca(NO_3)_2$ has a negligible effect on the total volume of solution.)

8. Given a 2.00 L saturated solution of $Cu(IO_3)_2$, $K_{sp} = 1.4 \times 10^{-7}$, how many moles of $NaIO_3$ would need to be dissolved in solution to reduce $[Cu^{2+}]$ to 6.0×10^{-5} M? (Assume the added $NaIO_3$ does not appreciably change the total volume of solution.)

Even more practice!

9. Copper(I) bromide has a measured solubility of 2.0×10^{-4} mol/L at 25°C. Calculate its K_{sp} value.

10. The K_{sp} value for copper(II) iodate, $Cu(IO_3)_2$, is 1.4×10^{-7} at 25°C. What is the maximum mass, in grams, of copper (II) iodate that can dissolve in 500. mL of water?

11. In pure water at 25°C, the molar solubility of $PbCl_2$ is 1.3×10^{-6} and the K_{sp} is 1.6×10^{-5}. LiCl is added to 5.00 L of a saturated solution of $PbCl_2$ at 25°C until the $[Pb^{2+}]$ is reduced to 4.5×10^{-4} M. How many moles of chloride ions are dissolved in solution at the point when $[Pb^{2+}] = 4.5 \times 10^{-4}$ M? (Assume the added LiCl has a negligible effect on the total volume of solution.)

Multiple Choice Practice!

Use the following information to answer questions 12−14.

150 mL of saturated SrF_2 solution is present in a 250 mL beaker at room temperature. The molar solubility of SrF_2 at 298 K is 1.0 x 10^{-3} M.

12. What are the concentrations of Sr^{2+} and F^- in the beaker?
    a. $[Sr^{2+}] = 1.0 \times 10^{-3}$ M; $[F^-] = 1.0 \times 10^{-3}$ M
    b. $[Sr^{2+}] = 1.0 \times 10^{-3}$ M; $[F^-] = 2.0 \times 10^{-3}$ M
    c. $[Sr^{2+}] = 2.0 \times 10^{-3}$ M; $[F^-] = 1.0 \times 10^{-3}$ M
    d. $[Sr^{2+}] = 2.0 \times 10^{-3}$ M; $[F^-] = 2.0 \times 10^{-3}$ M

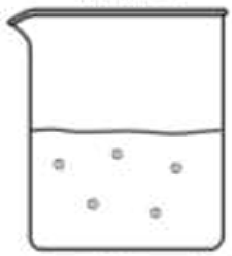
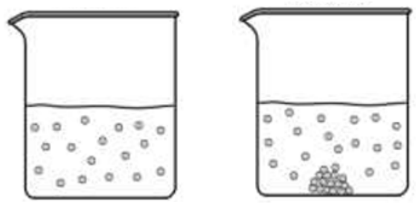
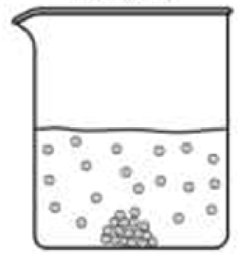
13. What would be the effect on $[Sr^{2+}]$ if some NaF(s) was added to the beaker?
    a. $[Sr^{2+}]$ would remain unchanged; neither ion in NaF(s) is common to Sr^{2+}
    b. $[Sr^{2+}]$ would increase; more Sr^{2+} ions would be needed to balance the additional F^- ions to re-establish equilibrium.
    c. $[Sr^{2+}]$ would decrease; the additional fluoride ions would cause the system to shift left to re-establish equilibrium.
    d. $[Sr^{2+}]$ would decrease; the additional Na^+ would cause an excess of positive charge, and the system would shift left to reduce overall positive charge.

14. Calculate the solubility product for SrF_2 at 25°C.
    a. 2 x 10^{-9}
    b. 2 x 10^{-6}
    c. 4 x 10^{-9}
    d. 4 x 10^{-6}

Will a Precipitate Form? A Task for K vs Q!

Precipitation occurs when the concentrations of ions is _________________ than the solubility of the ionic compound.

Compare the value of Q with given K_{sp} to determine if a precipitate will form!

K > Q	**K = Q**	**K < Q**
Unsaturated Solution	Saturated Solution	Saturated Solution with extra
System will shift right to reach equilibrium $\mathrm{AgCl(s)} \rightleftarrows \mathrm{Ag^+}(aq) + \mathrm{Cl^-}(aq)$	At equilibrium $\mathrm{AgCl(s)} \rightleftarrows \mathrm{Ag^+}(aq) + \mathrm{Cl^-}(aq)$	System will shift left to reach equilibrium $\mathrm{AgCl(s)} \rightleftarrows \mathrm{Ag^+}(aq) + \mathrm{Cl^-}(aq)$
More solid will dissolve until K = Q.	Solid will both dissolve and precipitate at the same rate.	**More solid will precipitate** until K = Q.
 no precipitate	 maybe precipitate	 yes precipitate

Important Ideas to Note:

1. If ______ solid is present, the solution is at equilibrium (a ____________________ solution)
2. Ion concentration, [ions], is **independent** of volume when at equilibrium (for instance, in a _________________ solution).
3. If ions are present that could form ______________ salts, the solid with the ______________ molar solubility will form.

Solubility Equilibrium Translation Guide

1. Solubility product constant = K_{sp} (aka the equilibrium constant for solubility)
2. Molar solubility = x from RICE table (aka how many moles of a solid will dissolve in 1.0 L, units = M = mol/L)
3. Saturated = equilibrium (aka a solution has dissolved as many ions as can fit, any extra will precipitate)

Let's Practice!

1. A chemist makes a 2.0 L saturated solution of $Ba_3(PO_4)_2$ solution, which has a K_{sp} = 6.0×10^{-39}.
   a. What is the concentration of Ba^{2+} ions in solution?

b. After two days of sitting on the counter, some liquid has evaporated from the solution. Did $[Ba^{2+}]$ increase, decrease, or remain the same? Justify your answer.

c. The chemist adds 3.00 g of solid $(NH_4)_3PO_4$ to the original saturated solution of $Ba_3(PO_4)_2$. Did $[Ba^{2+}]$ increase, decrease, or remain the same? Justify your answer.

2. A solution containing lead (II) nitrate is mixed with one containing sodium bromide to form a solution that is 0.0150 M in $Pb(NO_3)_3$ and 0.00350 M NaBr. Does a precipitate form in this newly mixed solution? (K_{sp} of $PbBr_2$ $= 4.67 \times 10^{-6}$)

3. The K_{sp} value for lead (II) bromide, $PbBr_2$, is 4.6×10^{-6} at 25°C. What is the maximum mass, in grams, of $PbBr_2$ that can dissolve in 1.50 L of water?

4. A student mixes 15.0 mL of 0.015 M sodium iodide solution, NaI, with 5.00 mL of 0.0025 M $Pb(NO_3)_2$. The K_{sp} of PbI_2 is 8.5 x 10^{-9} M. What will the student observe? Justify your answer with calculations.

5. Sodium carbonate is added to a 0.0024 M solution of the nickel (II) ion. If $[Na_2CO_3]$ = 1.0 x 10^{-4} M, will a precipitate form? (The K_{sp} of nickel (II) carbonate is 6.6 x 10^{-9}.)

6. Calculate the molar solubility of $Ba_3(PO_4)_2$, which has a $K_{sp} = 6.0 \times 10^{-39}$.

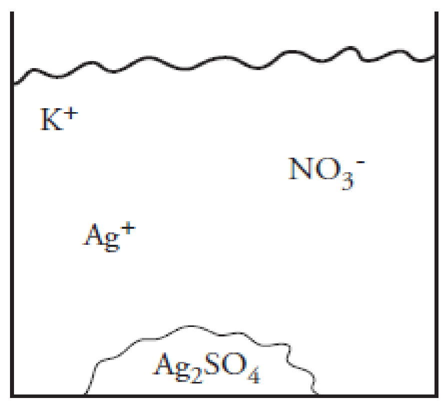
Multiple Choice Practice!

7. 150 mL of saturated SrF_2 solution is present in a 250 mL beaker at room temperature. If some of the solution evaporates overnight, which of the following will occur?
   a. The mass of the solid and the concentration of the ions will remain the same.
   b. The mass of the solid and the concentration of the ions will increase.
   c. The mass of the solid will decrease, and the concentration of the ions will remain the same.
   d. The mass of the solid will increase, and the concentration of the ions will remain the same.

8. A student added 1 liter of a 1.0 M KCl solution to 1 liter of a 1.0 M $Pb(NO_3)_3$ solution. A lead chloride precipitate formed, and nearly all of the lead ions disappeared from solution. Which of the following lists the ions remaining in the solution in order of decreasing concentration?
   a. $[NO_3^-] > [K^+] > [Pb^{2+}]$
   b. $[NO_3^-] > [Pb^{2+}] > [K^+]$
   c. $[K^+] > [Pb^{2+}] > [NO_3^-]$
   d. $[K^+] > [NO_3^-] > [Pb^{2+}]$

Use the following information to answer questions 8−10.

Silver sulfate, Ag_2SO_4, has a solubility product constant of 1.0×10^{-5}. The diagram to the right shows the products of a precipitation reaction in which some silver sulfate was formed.



9. What is the identity of the excess reactant?
   a. $AgNO_3$
   b. Ag_2SO_4
   c. $NaNO_3$
   d. Na_2SO_4

10. If the beaker above was left uncovered for several hours:
    I. Some of the Ag_2SO_4 would dissolve.
    II. Additional Ag_2SO_4 would precipitate.
    III. $[Ag^+]$ would remain constant.

    a. I only
    b. II only
    c. II and III
    d. I and III

11. Which ion concentration below would have led the precipitate to form?
    a. $[Ag^+] = 0.01$ M, $[SO_4^{2-}] = 0.01$ M
    b. $[Ag^+] = 0.10$ M, $[SO_4^{2-}] = 0.01$ M
    c. $[Ag^+] = 0.01$ M, $[SO_4^{2-}] = 0.10$ M
    d. It is impossible to determine without knowing the total volume of the solution.

Solubility Summary Sheet

Solubility Language	Normal Equilibrium Language
solubility product constant (K_{sp})	equilibrium constant (K)
molar solubility	x (from RICE table)
saturated solution	system at equilibrium

When you need to solve for molar solubility

Given?	Asked to find?	Strategy
K_{sp}	molar solubility	1. Write K_{sp} expression using ions produced when solid dissolves. 2. Substitute x values from mini-RICE table 3. Solve for x
1. K_{sp} OR 2. Concentration of all ions in a saturated solution	# of grams that can dissolve	1. Write K_{sp} expression using ions produced when solid dissolves. 2. Substitute x values from mini-RICE table 3. Solve for x (in M = mol / L) 4. Use the volume of solution to calculate moles that can dissolve 5. Use molar mass to convert to grams

When you don't need to solve for molar solubility

Given?	Asked to find?	Use:
Concentration of all ions in a saturated solution	K_{sp}	1. Write K_{sp} expression using ions produced when solid dissolves. 2. Substitute given concentrations and solve for K_{sp}.
other ion concentrations at equilibrium <u>and</u> K_{sp}	Concentration of ONE ion in a saturated solution	1. Write K_{sp} expression using ions produced when solid dissolves. 2. Plug in known values, solve for unknown concentration.
ion concentrations when solutions are added or mixed <u>and</u> K_{sp}	If a precipitate will form	1. Write Q_{sp} expression using ions produced when solid dissolves. 2. Substitute given concentrations and solve for Q_{sp}. 3. Compare K and Q: • K > Q = no precipitate • K < Q = yes precipitate

Unit 7 MC Practice

1.

$$2A(g) \rightleftharpoons 2\,B(g) + C(g) \qquad K = 1.6 \times 10^4$$

Two moles of Gas A are placed into a closed system where the temperature is held constant at 270 K and allowed to reach equilibrium as represented by the chemical reaction shown above. After 15 minutes at equilibrium, additional Gas B is injected into the reaction vessel. Which of the following best describes the behavior of the equilibrium mixture in response to the addition of Gas B?

A The concentration of Gas A increases.

B The concentration of Gas C increases.

C The value of the equilibrium constant increases.

D There is no observable effect since the equilibrium was already established.

2. In which of the following systems would the number of moles of the substances present at equilibrium NOT be shifted by a change in the volume of the system at constant temperature?

A $CO(g) + NO(g) \rightleftharpoons CO_2(g) + \frac{1}{2} N_2(g)$

B $N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$

C $N_2(g) + 2\,O_2(g) \rightleftharpoons 2\,NO_2(g)$

D $NO(g) + O_3(g) \rightleftharpoons NO_2(g) + O_2(g)$

3.

$$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$$

The chemical equation for the formation of hydrogen chloride gas from its elements is shown above. Given that the equilibrium constant for the reaction above is K_p, Which of the following best represents the equilibrium constant for the reaction

$$HCl(g) \rightleftharpoons \frac{1}{2}H_2(g) + \frac{1}{2}Cl_2(g)\ ?$$

A $\dfrac{1}{K_p^2}$

C $\dfrac{1}{\sqrt{K_p}}$

B K_p^2

D $\sqrt{K_p}$

4.

$$CaCl_2(s) + 2H_2O(g) \rightleftharpoons CaCl_2 \cdot 2H_2O(s)$$

Solid calcium chloride reacts with water as shown above. Which of the following is the correct equilibrium expression for this reaction?

A $K = \dfrac{[CaCl_2 \cdot 2H_2O]}{[CaCl_2][H_2O]}$

C $K = \dfrac{1}{2[H_2O]}$

B $K = \dfrac{1}{[H_2O]^2}$

D $K = [H_2O]^2$

5. A gaseous reaction occurs and comes to equilibrium, as shown below. Which of the following changes to the system will serve to increase the number of moles of I_2 present at equilibrium?

$$2\ HI(g) + Cl_2(g) \leftrightarrow 2\ HCl(g) + I_2(g) + \text{energy}$$

a. Increasing the volume at constant temperature

b. Decreasing the volume at constant temperature

c. Increasing the temperature at constant volume

d. Decreasing the temperature at constant volume

$$CO_2\ (g) + H_2\ (g) \rightleftharpoons CO\ (g) + H_2O\ (g) \quad K_c = 1.4$$
$$CO\ (g) + 2\ H_2\ (g) \rightleftharpoons CH_3OH\ (g) \quad K_c = 14.5$$

6. Given the above information, what would be the equilibrium constant for the reaction below?

$$3\ CO\ (g) + 2\ H_2O\ (g) \rightleftharpoons 2\ CO_2\ (g) + CH_3OH\ (g)$$

(A) $(2)(1.4)(14.5)$

(C) $\dfrac{14.5}{(1.4)^2}$

(B) $\dfrac{(1.4)(14.5)}{2}$

(D) $14.5 - 1.4^2$

7. In which of the following ways could the reaction below be manipulated to create more product?

$$PCl_3(g) + Cl_2(g) \leftrightarrow PCl_5(g) \qquad \Delta H = -92.5\ kJ/mol_{rxn}$$

a. Decreasing the concentration of PCl_3

b. Increasing the pressure

c. Increasing the temperature

d. None of the above

8. A sample of H_2S gas is placed in an evacuated, sealed container and heated until the following decomposition reaction occurs at 1000 K: $2\ H_2S(g) \rightarrow 2\ H_2(g) + S_2(g)$ $K_c = 1.0 \times 10^{-6}$

Which option best describes what will immediately occur to the reaction rates if the pressure on the system is increased after it has reached equilibrium?

a. The rate of both the forward and the reverse reactions will increase.

b. The rate of the forward reaction will increase will the rate of the reverse reaction will decrease.

c. The rate of the forward reaction will decrease will the rate of the reverse reaction will increase.

d. Neither the rate of the forward nor reverse reactions will change.

$$Br_2(g) + I_2(g) \leftrightarrow 2\ IBr(g)$$

9. At 150°C, the equilibrium constant, K_p, for the reaction shown above has a value of 300. The reaction was allowed to reach equilibrium in a sealed container and the partial pressure due to IBr(g) was found to be 3 atm. Which of the following could be the partial pressures due to $Br_2(g)$ and $I_2(g)$ in the container?

	$Br_2(g)$	$I_2(g)$
a.	0.1 atm	0.3 atm
b.	0.3 atm	1 atm
c.	1 atm	1 atm
d.	1 atm	3 atm

10. Consider the following equilibrium: $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ $K_{eq} = 50.0$

What is the value K_{eq} for the reaction rewritten as: $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$ $K_{eq} = ?$

a. −50.0 b. 0.0200 c. 25.0 d. 50.0

11. Consider the following equilibrium: $N_2O_4(g) \rightleftharpoons 2NO_2(g)$

An equilibrium mixture contains 4.0×10^{-2} mol N_2O_4 and 1.5×10^{-2} mol NO_2 in a 1.0 L flask. What is the value of K_{eq}?

a. 5.6×10^{-3} b. 3.8×10^{-1} c. 7.5×10^{-1} d. 1.8×10^{2}

12. Consider the following equilibrium: $Cl_2(g) + 2\ NO(g) \rightleftharpoons 2\ NOCl(g)$ $K_{eq} = 5.0$

At equilibrium, $[Cl_2] = 1.0$ M and $[NO] = 2.0$ M. What is the [NOCl] at equilibrium?

a. 0.12 M b. 0.89 M c. 4.5 M d. 10. M

AP Free Response Practice #1 [2007 #1, 6 points, modified]

1. A sample of solid U_3O_8 is placed in a rigid 1.500 L flask. Chlorine gas, $Cl_2(g)$, is added, and the flask is heated to 862°C. The equation for the reaction that takes place and the equilibrium constant expression for the reaction are given below.

$$\mathrm{U_3O_8}(s) + 3\ \mathrm{Cl_2}(g) \leftrightarrow 3\ \mathrm{UO_2Cl_2}(g) + \mathrm{O_2}(g)$$

When the system is at equilibrium, the partial pressure of $Cl_2(g)$ is 1.007 atm, the partial pressure of $UO_2Cl_2(g)$ is 9.734×10^{-4} atm, and the partial pressure of $O_2(g)$ is 3.245×10^{-4} atm.

a. Write the equilibrium constant expression, K_p, for this reaction and calculate the value of K_p for the system at 862°C. [2 points]

b. Calculate the concentration of $Cl_2(g)$ at equilibrium. [1 point]

c. Calculate the Gibb's free-energy change, ΔG°, for the reaction at 862°C. [1 point]

d. After a certain period of time, 1.000 mole of $O_2(g)$ is added to the mixture in the flask. Does the mass of $U_3O_8(s)$ in the flask increase, decrease, or remain the same? Justify your answer. [2 point]

AP Free Response Practice #2 [2010, 10 points]

2. Several reactions are carried out using AgBr, a cream-colored silver salt for which the value of the solubility-product constant, K_{sp}, is 5.0×10^{-13} at 298 K.

   a. Write the expression for the solubility-product constant, K_{sp}, of AgBr. [1 point]

   b. Calculate the value of $[Ag^+]$ in 50.0 mL of a saturated solution of AgBr at 298 K. [1 point]

   c. A 50.0 mL sample of distilled water is added to the solution described in part (b), which is in a beaker with some solid AgBr at the bottom. The solution is stirred and equilibrium is reestablished. Some solid AgBr remains in the beaker. Is the value of $[Ag^+]$ greater than, less than, or equal to the value you calculated in part (b) ? Justify your answer. [1 point]

   d. Calculate the minimum volume of distilled water, in liters, necessary to completely dissolve a 5.0 g sample of AgBr(*s*) at 298 K. (The molar mass of AgBr is 188 g mol^{-1}.) [2 points]

e. A student mixes 10.0 mL of 1.5×10^{-4} *M* $AgNO_3$ with 2.0 mL of 5.0×10^{-4} *M* NaBr and stirs the resulting mixture. What will the student observe? Justify your answer with calculations. [3 points]

f. The color of another salt of silver, AgI(*s*), is yellow. A student adds a solution of NaI to a test tube containing a small amount of solid, cream-colored AgBr. After stirring the contents of the test tube, the student observes that the solid in the test tube changes color from cream to yellow.

i. Write the chemical equation for the reaction that occurred in the test tube. [1 point]

ii. Which salt has the greater value of K_{sp}: AgBr or AgI ? Justify your answer. [1 point]

AP Free Response Practice #3 [2011B #1, modified, 9 points]

3. Answer the following questions about the solubility and reactions of the ionic compounds $M(OH)_2$ and MCO_3, where M represents an unidentified metal.

   a. Identify the charge of the M ion in the ionic compounds above. [1 point] ________________

   b. At 25°C, in a saturated solution of $M(OH)_2$, $[OH^-] = 1.4 \times 10^{-5}$ M.

      i. Write the solubility-product constant expression for $M(OH)_2$. [1 point]

      ii. Calculate the value of the solubility-product constant, K_{sp}, for $M(OH)_2$ at 25°C. [2 points]

   c. For the metal carbonate, MCO_3, the value of the solubility-product constant, K_{sp}, is 7.4×10^{-14} at 25°C. On the basis of this information and your results in part (b), which compound, $M(OH)_2$ or MCO_3, has the greater molar solubility in water at 25°C? Justify your answer with a calculation. [2 points]

d. MCO_3 decomposes at high temperatures, as shown by the reaction represented below.

$$MCO_3(s) \leftrightarrow MO(s) + CO_2(g)$$

A sample of MCO_3 is placed in a previously evacuated container, heated to 423 K, and allowed to come to equilibrium. Some solid MCO_3 remains in the container. The value of K_p for the reaction at 423 K is 0.0012.

i. Write the equilibrium-constant expression for K_p of the reaction. [1 point]

ii. Determine the pressure, in atm, of $CO_2(g)$ in the container at equilibrium at 423 K. [1 point]

iii. Indicate whether the value of $\Delta G°$ for the reaction at 423 K is positive, negative, or zero. Justify your answer.

AP Free Response Practice #4 [Laying the Foundation, 9 points]

4. When 3.29 moles of nitrogen gas, N_2, and 2.95 moles of hydrogen gas, H_2, are placed in a 5.00 L evacuated container at 773 K, ammonia gas, NH_3, is formed according to this equation: $N_2(g) + 3\ H_2(g) \leftrightarrow 2\ NH_3(g)$

   When equilibrium is established, 0.565 mole of $NH_3(g)$ is present in the flask.

   a. Write the expression for the equilibrium constant, K_c, for this reaction. [1 point]

   b. Calculate the equilibrium concentrations, in mol L^{-1}, of the following gases in the container at 773 K.

      i. $N_2(g)$ [1 point]

      ii. $H_2(g)$ [1 point]

c. Calculate the value of the equilibrium constant, K_c, at 773 K. [1 point]

d. When the same reaction is carried out at 298 K, the number of moles of NH_3 present at equilibrium is much larger than 0.565 mole. Is the forward reaction endothermic or exothermic? Justify your answer. [1 point]

e. Calculate the value of the equilibrium constant, K_c, for the reaction below. [1 points]

$$6\ NH_3(g) \leftrightarrow 3\ N_2(g) + 9\ H_2(g)$$

AP Free Response Practice #5 [5 Steps to a 5, 7 points]

$$N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g) \qquad \Delta H = -92.4\ \text{kJ/mol}_{rxn}$$

1. When the reaction above took place at a temperature of 570 K in a reaction vessel of 0.75 L, the following values were measured after the system reached equilibrium: NH_3 = 0.150 mol, N_2 = 0.375 mol, and H_2 = 0.150 mol.

   a. Write the expression for the equilibrium constant, K_c, and calculate its value. [2 points]

   b. Calculate ΔG for this reaction. [1 point]

   c. Describe how the concentration of $H_2(g)$ at equilibrium will be affected by each of the following changes to the system at equilibrium.

      i. The temperature is increased. [1 point]

ii. The volume of the reaction chamber is increased. [1 point]

iii. N_2 gas is added to the reaction chamber. [1 point]

iv. Helium gas is added to the reaction chamber. [1 point]

AP Free Response Practice #6 [2013 #1, modified, 4 points]

2. Answer the following questions about the solubility of some fluoride salts of alkaline earth metals.

   a. A student prepares 100. mL of a saturated solution of MgF_2 by adding 0.50 g of solid MgF_2 to 100. mL of distilled water at 25°C and stirring until no more solid dissolves. (Assume that the volume of the undissolved MgF_2 is negligibly small.) The saturated solution is analyzed, and it is determined that $[F^-]$ in the solution is 2.4×10^{-3} *M*.

      i. Write the chemical equation for the dissolving of solid MgF_2 in water. [1 point]

      ii. Calculate the number of moles of MgF_2 that dissolved. [2 points]

      iii. Determine the value of the solubility-product constant, K_{sp}, for MgF_2 at 25°C. [1 point]

AP Free Response Practice #7 [2015, 4 points]

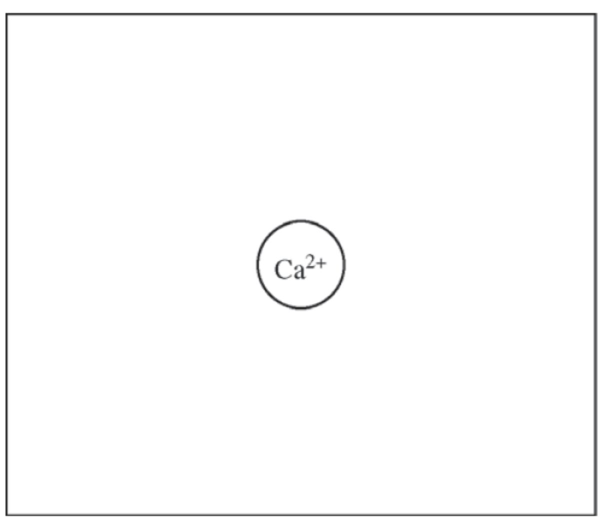
5. Answer the following questions about the solubility of $Ca(OH)_2$ ($K_{sp} = 1.3 \times 10^{-6}$). (2015)

   b. Write a balanced chemical equation for the dissolution of $Ca(OH)_2$ in pure water. [1 point]

   c. Calculate the molar solubility of $Ca(OH)_2$ in 0.10 M $Ca(NO_3)_2$. [2 points]

   d. In the box below, complete a particle representation diagram that includes <u>four</u> water molecules with proper orientation around the Ca^{2+} ion. [1 point]

   Represent water molecules as .



Polyatomic Ions

1st six-weeks

Nick the **C**amel ***ate*** an **I**cky **Cl**am for **S**upper in **P**hoenix with his **Br**os

NO_3^-	nitrate	ClO_3^-	chlorate	PO_3^{3-}	phosphite
NO_2^-	nitrite	ClO_2^-	chlorite	BrO_3^-	bromate
CO_3^{2-}	carbonate	SO_4^{2-}	sulfate	BrO_2^-	bromite
IO_3^-	iodate	SO_3^{2-}	sulfite		
IO_2^-	iodite	PO_4^{3-}	phosphate		

Consonants = # of Oxygen

Vowels = Charge

2nd six-weeks

ClO_4^-	perchlorate	IO^-	hypoiodite
ClO^-	hypochlorite	BrO_4^-	perbromate
IO_4^-	periodate	BrO^-	hypobromite

	Difference in Oxygen from ATE
Per_____ate	+1
Ate	0
Ite	−1
Hypo ____ite	−2

3rd six-weeks

$H_2PO_4^-$	dihydrogen phosphate	HCO_3^-	hydrogen carbonate *or* bicarbonate
HPO_4^{2-}	hydrogen phosphate	HSO_4^-	hydrogen sulfate

4th six-weeks

NH_4^+	ammonium	OH^-	hydroxide
$C_2H_3O_2^-$ *or* CH_3COO^-	acetate	H_3O^+	hydronium

5th six-weeks

MnO_4^-	permanganate	CrO_4^{2-}	chromate
CN^-	cyanide	$Cr_2O_7^{2-}$	dichromate

6th six-weeks

O_2^{2-}	peroxide	$C_4H_4O_6^{2-}$	tartrate
$S_2O_3^{2-}$	thiosulfate	$C_2O_4^{2-}$	oxalate

Acid Nomenclature	
Binary	Hydro______ic
ate	ic
ite	ous

Polyatomic Ions

Br–Based Ions

BrO^-	hypobromite
BrO_2^-	bromite
BrO_3^-	bromate
BrO_4^-	perbromate

Cr–Based Ions

CrO_4^{2-}	chromate
$Cr_2O_7^{2-}$	dichromate

I–Based Ions

IO^-	hypoiodite
IO_2^-	iodite
IO_3^-	iodate
IO_4^-	periodate

P–Based Ions

PO_3^{3-}	phosphite
PO_4^{3-}	phosphate
HPO_4^{2-}	hydrogen phosphate
$H_2PO_4^-$	dihydrogen phosphate

Other Ions

CN^-	cyanide
O_2^{2-}	peroxide
MnO_4^-	permanganate

C–Based Ions

CO_3^{2-}	carbonate
HCO_3^-	hydrogen carbonate *or* bicarbonate
$C_2H_3O_2^-$ *or* CH_3COO^-	acetate
$C_4H_4O_6^{2-}$	tartrate
$C_2O_4^{2-}$	oxalate

Cl–Based Ions

ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate

N–Based Ions

NO_2^-	nitrite
NO_3^-	nitrate
NH_4^+	ammonium

S–Based Ions

SO_3^{2-}	sulfite
SO_4^{2-}	sulfate
HSO_4^-	hydrogen sulfate
$S_2O_3^{2-}$	thiosulfate

Acid & Base Ions

H_3O^+	hydronium
OH^-	hydroxide

PERIODIC TABLE OF THE ELEMENTS

Atomic number — 14
Symbol — Si
Atomic mass — 28.09
Silicon — Name

	1	2											3	4	5	6	7	8
1	1 **H** 1.008 Hydrogen																	2 **He** 4.00 Helium
2	3 **Li** 6.94 Lithium	4 **Be** 9.01 Beryllium											5 **B** 10.81 Boron	6 **C** 12.01 Carbon	7 **N** 14.01 Nitrogen	8 **O** 16.00 Oxygen	9 **F** 19.00 Fluorine	10 **Ne** 20.18 Neon
3	11 **Na** 22.99 Sodium	12 **Mg** 24.30 Magnesium											13 **Al** 26.98 Aluminum	14 **Si** 28.09 Silicon	15 **P** 30.97 Phosphorus	16 **S** 32.06 Sulfur	17 **Cl** 35.45 Chlorine	18 **Ar** 39.95 Argon
4	19 **K** 39.10 Potassium	20 **Ca** 40.08 Calcium	21 **Sc** 44.96 Scandium	22 **Ti** 47.90 Titanium	23 **V** 50.94 Vanadium	24 **Cr** 52.00 Chromium	25 **Mn** 54.94 Manganese	26 **Fe** 55.85 Iron	27 **Co** 58.93 Cobalt	28 **Ni** 58.69 Nickel	29 **Cu** 63.55 Copper	30 **Zn** 65.39 Zinc	31 **Ga** 69.72 Gallium	32 **Ge** 72.59 Germanium	33 **As** 74.92 Arsenic	34 **Se** 78.96 Selenium	35 **Br** 79.90 Bromine	36 **Kr** 83.80 Krypton
5	37 **Rb** 85.47 Rubidium	38 **Sr** 87.62 Strontium	39 **Y** 88.91 Yttrium	40 **Zr** 91.22 Zirconium	41 **Nb** 92.91 Niobium	42 **Mo** 95.94 Molybdenum	43 **Tc** (98) Technetium	44 **Ru** 101.10 Ruthenium	45 **Rh** 102.91 Rhodium	46 **Pd** 106.42 Palladium	47 **Ag** 107.87 Silver	48 **Cd** 112.41 Cadmium	49 **In** 114.82 Indium	50 **Sn** 118.71 Tin	51 **Sb** 121.75 Antimony	52 **Te** 127.60 Tellurium	53 **I** 126.91 Iodine	54 **Xe** 131.29 Xenon
6	55 **Cs** 132.91 Cesium	56 **Ba** 137.33 Barium	57 ***La** 138.91 Lanthanum	72 **Hf** 178.49 Hafnium	73 **Ta** 180.95 Tantalum	74 **W** 183.85 Tungsten	75 **Re** 186.21 Rhenium	76 **Os** 190.2 Osmium	77 **Ir** 192.2 Iridium	78 **Pt** 195.08 Platinum	79 **Au** 196.97 Gold	80 **Hg** 200.59 Mercury	81 **Tl** 204.38 Thallium	82 **Pb** 207.2 Lead	83 **Bi** 208.98 Bismuth	84 **Po** (209) Polonium	85 **At** (210) Astatine	86 **Rn** (222) Radon
7	87 **Fr** (223) Francium	88 **Ra** 226.02 Radium	89 **†Ac** 227.03 Actinium	104 **Rf** (261) Rutherfordium	105 **Db** (262) Dubnium	106 **Sg** (266) Seaborgium	107 **Bh** (264) Bohrium	108 **Hs** (277) Hassium	109 **Mt** (268) Meitnerium	110 **Ds** (271) Darmstadtium	111 **Rg** (272) Roentgenium							

***Lanthanide Series**	58 **Ce** 140.12 Cerium	59 **Pr** 140.91 Praseodymium	60 **Nd** 144.24 Neodymium	61 **Pm** (145) Promethium	62 **Sm** 150.4 Samarium	63 **Eu** 151.97 Europium	64 **Gd** 157.25 Gadolinium	65 **Tb** 158.93 Terbium	66 **Dy** 162.5 Dysprosium	67 **Ho** 164.93 Holmium	68 **Er** 167.26 Erbium	69 **Tm** 168.93 Thulium	70 **Yb** 173.04 Ytterbium	71 **Lu** 174.97 Lutetium
†Actinide Series	90 **Th** 232.04 Thorium	91 **Pa** 231.04 Protactnium	92 **U** 238.03 Uranium	93 **Np** (237) Neptumium	94 **Pu** (244) Plutonium	95 **Am** (243) Americium	96 **Cm** (247) Curium	97 **Bk** (247) Berkelium	98 **Cf** (251) Californium	99 **Es** (252) Einsteinium	100 **Fm** (257) Fermium	101 **Md** (258) Mendelevium	102 **No** (259) Nobelium	103 **Lr** (262) Lawrencium

AP® CHEMISTRY EQUATIONS AND CONSTANTS

Throughout the exam the following symbols have the definitions specified unless otherwise noted.

L, mL	=	liter(s), milliliter(s)	mm Hg	=	millimeters of mercury
g	=	gram(s)	J, kJ	=	joule(s), kilojoule(s)
nm	=	nanometer(s)	V	=	volt(s)
atm	=	atmosphere(s)	mol	=	mole(s)

ATOMIC STRUCTURE

$$E = h\nu$$
$$c = \lambda\nu$$

E = energy
ν = frequency
λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s
Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
Avogadro's number $= 6.022 \times 10^{23}$ mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}, \text{ where } a\,A + b\,B \rightleftarrows c\,C + d\,D$$

$$K_p = \frac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$K_b = \frac{[OH^-][HB^+]}{[B]}$$

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$
$$= K_a \times K_b$$

$$\text{pH} = -\log[H^+],\ \text{pOH} = -\log[OH^-]$$

$$14 = \text{pH} + \text{pOH}$$

$$\text{pH} = \text{p}K_a + \log\frac{[A^-]}{[HA]}$$

$$\text{p}K_a = -\log K_a,\ \text{p}K_b = -\log K_b$$

Equilibrium Constants

K_c (molar concentrations)
K_p (gas pressures)
K_a (weak acid)
K_b (weak base)
K_w (water)

KINETICS

$$\ln[A]_t - \ln[A]_0 = -kt$$

$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$$

$$t_{1/2} = \frac{0.693}{k}$$

k = rate constant
t = time
$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{\text{total}} \times X_{\text{A}}, \text{ where } X_{\text{A}} = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_{\text{A}} + P_{\text{B}} + P_{\text{C}} + \ldots$$

$$n = \frac{m}{\boldsymbol{M}}$$

$$\text{K} = {}^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

$$\text{Molarity, } M = \text{moles of solute per liter of solution}$$

$$A = abc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
$\boldsymbol{M}$ = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
a = molar absorptivity
b = path length
c = concentration

Gas constant, R = 8.314 J mol^{-1} K^{-1}
= 0.08206 L atm mol^{-1} K^{-1}
= 62.36 L torr mol^{-1} K^{-1}
1 atm = 760 mm Hg = 760 torr
STP = 273.15 K and 1.0 atm
Ideal gas at STP = 22.4 L mol^{-1}

THERMODYNAMICS / ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\begin{aligned} \Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ \\ &= -RT \ln K \\ &= -nFE^\circ \end{aligned}$$

$$I = \frac{q}{t}$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)

Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

PERIODIC TABLE OF THE ELEMENTS

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.10	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.91	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 ***La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.21	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.02	89 **†Ac** 227.03	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (277)	109 **Mt** (268)	110 **Ds** (271)	111 **Rg** (272)							

***Lanthanide Series**	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.5	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
†Actinide Series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)