Recitation Worksheet Eight

Name: _____ *Key*   MyID: _____

Textbook:
Chemistry & Chemical Reactivity
by John C. Kotz, Paul M. Treichel, John R. Townsend, David Treichel
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Instructions:
- This recitation worksheet covers Ch. 15.1-15.6, 18.5-18.7
- Please enter your first and last name as it <u>appears on the eLC roster</u> (do not use a nickname that is not reflected in eLC).
- Your UGA myID is a combination of letters and numbers (example: Dr. Abdelrahman's MyID is ema88805@uga.edu). **Do not use your 81x number**.
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    - You may use an app to annotate the worksheet by placing your answers in the answer boxes and showing your work when appropriate. Afterward, submit the worksheet to Gradescope. You will not need to upload anything to eLC.
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- The following criteria **must** be met to be eligible for full credit:
    - You must make sure the pages are in the <u>correct order</u> and have the <u>same layout</u> as the original worksheet when submitting to Gradescope regardless of your submission type.
    - Answers must be written in the corresponding answer boxes.
    - You must show your work when appropriate.
- This worksheet is due no later than **12:00 PM (noon) on the Saturday, October 19th.**
- A periodic table and formula sheet are attached to the end of this worksheet. Please keep these attached to your worksheet in the correct order when submitting to Gradescope.

1. Use the equation below to answer the following questions:

$$Fe(OH)_3(s) + 3\ H^+(aq) \rightleftharpoons Fe^{3+}(aq) + 3\ H_2O(l) \quad K = 9.1 \times 10^3$$
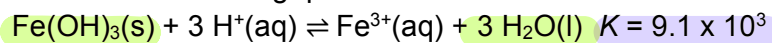
solids & liquids are not included in the equilibrium constant expression

A. The correct equilibrium constant expression for this reaction is

iv

i. $\dfrac{[FeOH_3][H^+]^3}{[Fe^{3+}][H_2O]^3}$

ii. $\dfrac{[Fe^{3+}][H_2O]^3}{[FeOH_3][H^+]^3}$

iii. $\dfrac{[Fe^{3+}][H_2O]}{[FeOH_3][H^+]}$

iv. $\dfrac{[Fe^{3+}]}{[H^+]^3}$

v. $\dfrac{[FeOH_3][H^+]}{[Fe^{3+}][H_2O]}$

$$K = \dfrac{[Fe^{3+}]}{[H^+]^3}$$

B. What is the equilibrium concentration for $[Fe^{3+}]$ if the equilibrium concentration for $[H^+] = 2.5 \times 10^{-8}$? Use **scientific notation** to report your answer.
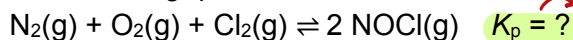
1.4 $\times 10^{\boxed{-19}}$ M

$$K = 9.1 \times 10^3$$

$$9.1 \times 10^3 = \dfrac{[Fe^{3+}]}{[2.5 \times 10^{-8}]^3}$$
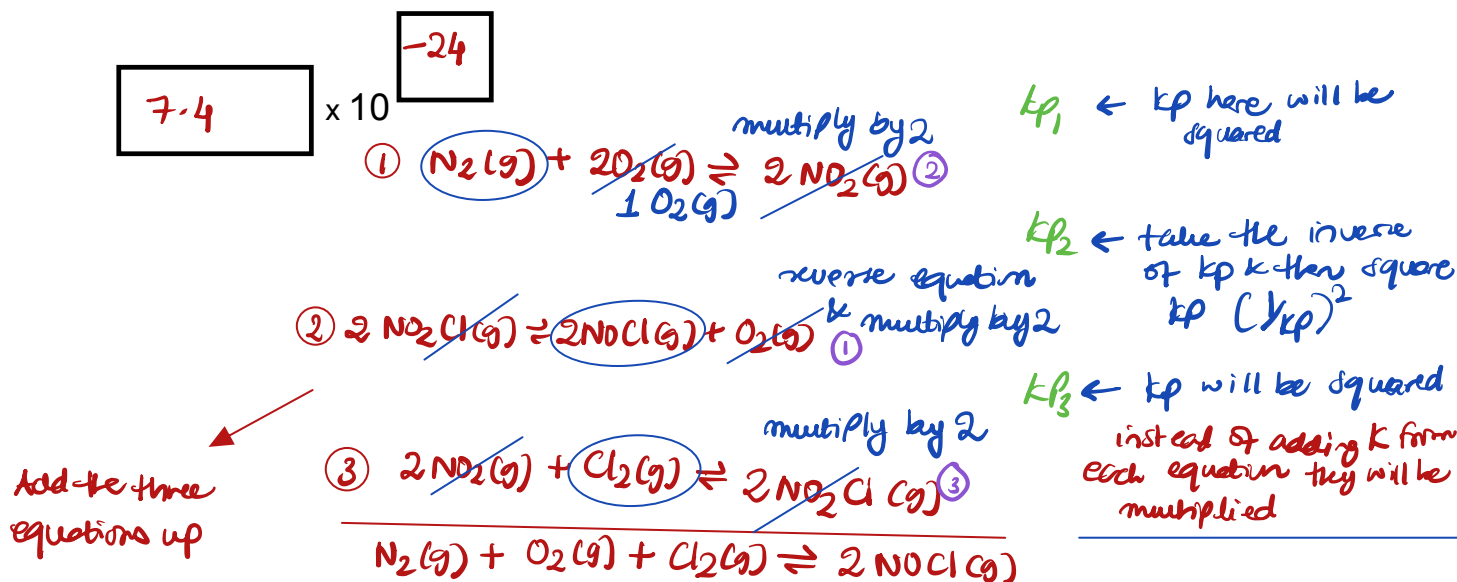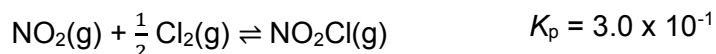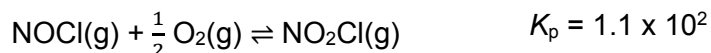
$$[Fe^{3+}] = 1.421876 \times 10^{-19}$$

$$\approx 1.4 \times 10^{-19}\ M$$

2. Use the reaction below to answer the following questions

$$N_2(g) + O_2(g) + Cl_2(g) \rightleftharpoons 2\,NOCl(g) \quad K_p = ?$$

Altering K does not follow the same rules as Hess's law

A. Calculate K_p at 25 °C using the set of equations provided. Use **scientific notation** to report your answer.

$\frac{1}{2} N_2(g) + O_2(g) \rightleftharpoons NO_2(g)$	$K_p = 1.0 \times 10^{-9}$
$NOCl(g) + \frac{1}{2} O_2(g) \rightleftharpoons NO_2Cl(g)$	$K_p = 1.1 \times 10^2$
$NO_2(g) + \frac{1}{2} Cl_2(g) \rightleftharpoons NO_2Cl(g)$	$K_p = 3.0 \times 10^{-1}$

$\boxed{7.4}$ x 10$^{\boxed{-24}}$

① $N_2(g) + 2O_2(g) \rightleftharpoons 2NO_2(g)$ ②   *multiply by 2* $1\,O_2(g)$

$Kp_1 \leftarrow$ *kp here will be squared*

② $2NO_2Cl(g) \rightleftharpoons 2NOCl(g) + O_2(g)$ ①   *reverse equation, be multiply by 2*

$Kp_2 \leftarrow$ *take the inverse of kp & then square kp* $\left(\frac{1}{kp}\right)^2$

Add the three equations up

③ $2NO_2(g) + Cl_2(g) \rightleftharpoons 2NO_2Cl(g)$ ③   *multiply by 2*

$Kp_3 \leftarrow$ *kp will be squared instead of adding k from each equation they will be multiplied*

$\overline{N_2(g) + O_2(g) + Cl_2(g) \rightleftharpoons 2\,NOCl(g)}$

$K_{tol} \quad Kp_1 \times Kp_2 \times Kp_3 =$

$(1.0 \times 10^{-9})^2 \times \left(\frac{1}{1.1 \times 10^2}\right)^2 \times (3.0 \times 10^{-1})^2$

$= 7.4 \times 10^{-24}$

B. Calculate K_c for the same reaction.

$\boxed{1.8}$ x 10$^{\boxed{-22}}$

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

$K_p = 7.4 \times 10^{-24}$

$R = 0.08204 \frac{L \cdot atm}{mol \cdot K}$

$T = 25 + 273.15 = 298K$

$\Delta n = 2 - 3 = 1$

$7.4 \times 10^{-24} = K_c \left(0.08206 \frac{L \cdot atm}{mol \cdot K} \times 298\right)^{(2-3)}$
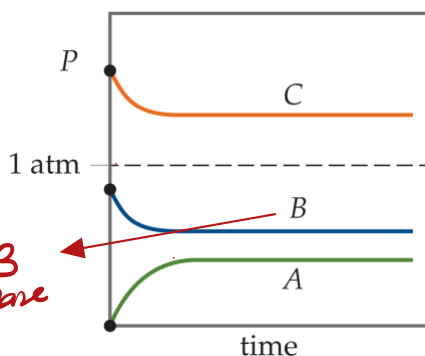
$K_c = 1.81888364 \times 10^{-22}$

$\sim 1.83 \times 10^{-22}$

$$K_p = \frac{(PC)}{(PA)(PB)}$$

3. For the reaction A(g) + B(g) ⇌ C(g), $K_p < 1$. Which of the following charts below describes the approach to equilibrium of a mixture of B(g) and C(g)?

→ Since $K_p < 1$ this means that the reaction is reactant favored &
∴ C will produce A & B ⌒

↓
the conc.
of C decreases
to form A & B

the conc. of A & B increases

B

A.

the conc of B should increase as well ←

P
C
1 atm
B
A
time

B.

$kp = \frac{(0.3)}{(1.1)(1.6)}$

$= 0.17$

$kp < 1$

P
1.6 — B
1.1 — A
1 atm
0.3 — C
time

---

e.

$kp = \frac{(1.5)}{(0.6)(0.4)}$

6.25

∴ $kp > 1$

P
1.5 — C
1 atm
0.6 — B
0.4 — A
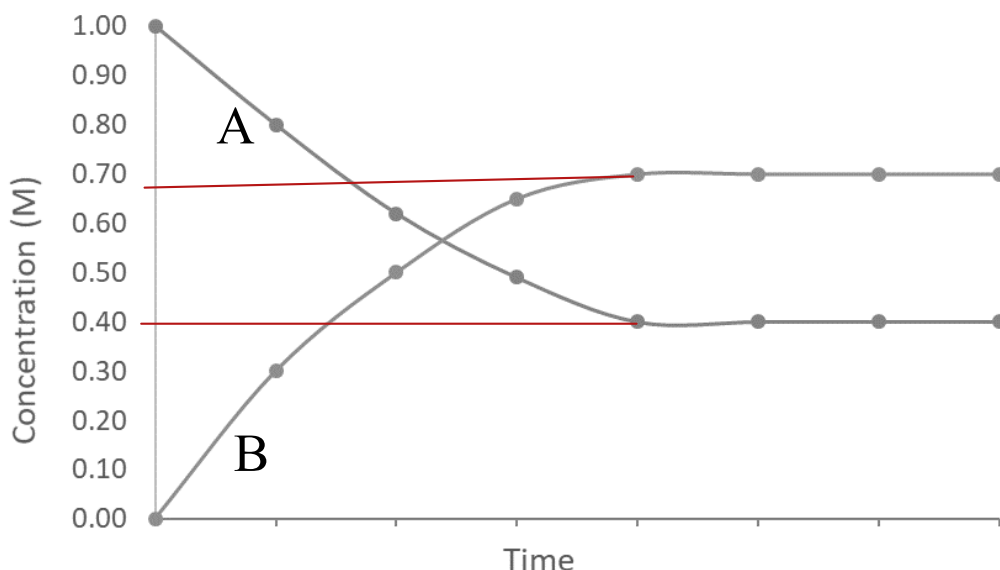time

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D.

P
A
1 atm — B
C
time
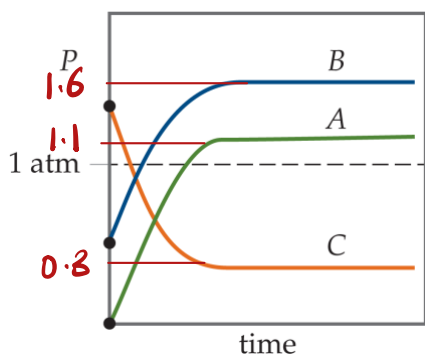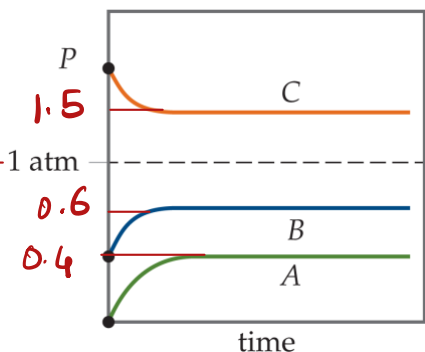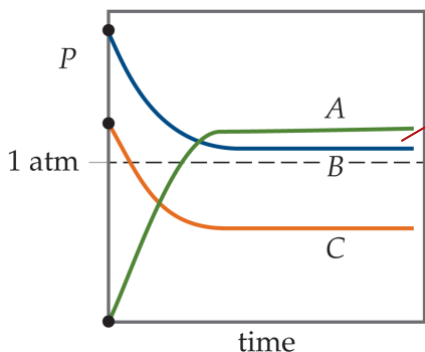
the conc. of B should increase as well →

4

4. A graph for the reaction $A(g) \rightleftharpoons 2B(g)$ shows the change in concentration over time. <mark>What is the equilibrium constant for this reaction?</mark>

B

A. 1.8
B. 1.2
C. 4.4
D. 0.82
E. 0.57

$K_C = \dfrac{[B]^2}{[A]}$

$= \dfrac{[0.70]^2}{[0.40]} = 1.225$

~ 1.2

Graph: Concentration (M) vs Time. Y-axis from 0.00 to 1.00. Curve A decreases from 1.00 to ~0.40. Curve B increases from 0.00 to ~0.70. Red horizontal lines at 0.70 and 0.40.

5. You prepare a <mark>saturated solution of PbBr₂</mark>, and the equilibrium equation is written as follows:

$\rightsquigarrow$ In a saturated solution the dissolved solute is in dynamic equilibrium with the undissolved solute

$$PbBr_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\,Br^-(aq)$$

If solid PbBr₂ is added to this solution, which of the following statements is **true**?

C

A. The concentration of Pb^{2+} will increase more than the concentration of Br-
B. The concentration of Br^- will increase more than the concentration of Pb^{2+}
C. The concentration of Pb^{2+} and Br^- will be the same as they were before more PbBr₂ was added
D. The concentration of Pb^{2+} and Br^- will increase after more PbBr₂ was added
E. The value of K_c will decrease

the addition of solute to a saturated solution will not increase the conc. of the ions in the solution

6. In which of the following reactions will <mark>$K_p = K_c$</mark> ? <u>Select all that apply.</u> Insert letters without spaces in the answer box, example **ABCD**.

BD

A. $4\,KO_2(s) + 2\,CO_2(g) \rightleftharpoons 2\,K_2CO_3(s) + 3\,O_2(g)$ $\Delta n = 3-2=1$
B. $3\,Fe(s) + 4\,H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4\,H_2(g)$ $\Delta n = 4-4=0$
C. $2\,H_2S(g) + CH_4(g) \rightleftharpoons CS_2(g) + 4\,H_2(g)$ $\Delta n = 5-3=2$
D. $2\,HI(g) \rightleftharpoons H_2(g) + I_2(g)$ $\Delta n = 2-2=0$
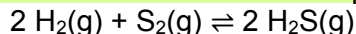E. $4\,NH_3(g) + 5\,O_2(g) \rightleftharpoons 4\,NO(g) + 6\,H_2O(g)$ $\Delta n = 10-9=1$

$K_p = K_c$ when $\Delta n = 0$

$K_p = K_c (RT)^{\Delta n}$
$K_p = K_c (RT)^0$
$\therefore K_p = K_c$

5

7. A mixture of 1.00 g H_2 and 1.06 g H_2S in a 0.500 L flask comes to equilibrium at 1670 K:
$$2 H_2(g) + S_2(g) \rightleftharpoons 2 H_2S(g)$$

The equilibrium amount of $S_2(g)$ found is 8.00 x 10^{-6} mol.

A. What is the equilibrium concentration of H_2S?

0.0622 M

$[H_2] = \dfrac{1.00g \times \frac{1mol}{2.02g}}{0.500 L} = 0.990099$ M $\quad [H_2S] = \dfrac{1.06g \times \frac{1mol}{34.10g}}{0.500 L} = 0.062170088$ M

the equilibrium concentration $\rightarrow [S_2] = \dfrac{8.00 \times 10^{-6} mol}{0.500 L} = 1.60 \times 10^{-5}$ M

to determine the direction of the reaction we are Q & K

I	0.990	0	0.0622
C	2(+1.60×10⁻⁵)	+1.60×10⁻⁵	2(-1.60×10⁻⁵)
E	0.990	+1.6×10⁻⁵	0.0622

$Q = \dfrac{[H_2S]^2}{[H_2]^2[S_2]}$

$= \dfrac{[0.0622]^2}{[0.990]^2[0]}$

$= \infty$

∴ the equilibrium reaction proceeds towards products

B. What is the value of K_c?

246

$K_c = \dfrac{[H_2S]^2}{[H_2]^2[S_2]}$

$= \dfrac{[0.0622]^2}{[0.990]^2[1.60 \times 10^{-5}]} = 246.4748$
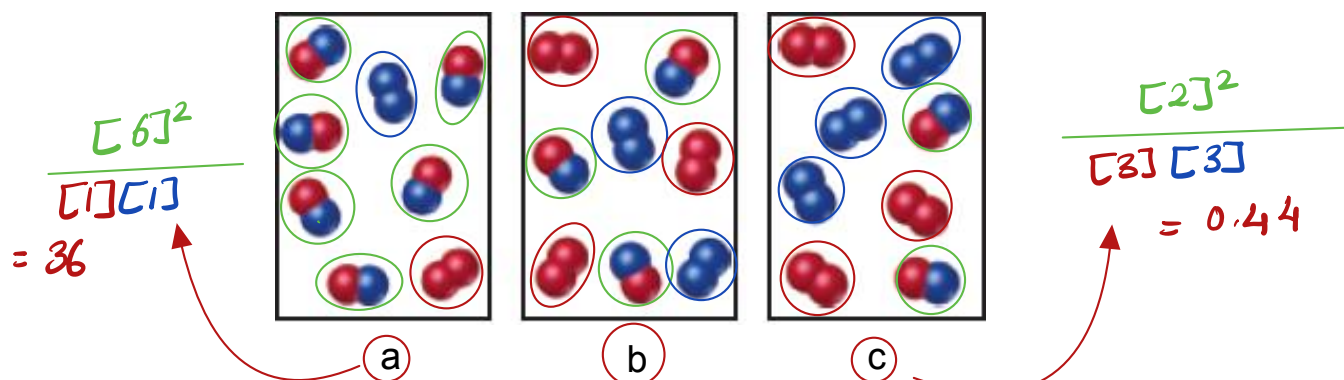~ 246

C. What is the value of K_p?

1.80

$K_p = K_c (RT)^{\Delta n}$
$= 246 (0.08206 \times 1670)^{2-3}$
$= 1.7950937$
~ 1.80

6

8. The diagram below represents the reaction between $A_2(g) + B_2(g) \rightleftharpoons 2\,AB(g)$, A<ins>$_2$ molecules are represented</ins> <ins>by red spheres, B$_2$ molecules are represented by blue spheres</ins> and the reaction has an equilibrium constant, $K_c = 1.5$. Which of the following diagrams a – c represents:

$$\frac{[6]^2}{[1][1]} = 36$$

$$\frac{[3]^2}{[3][2]} = 1.5$$

$$\frac{[2]^2}{[3][3]} = 0.44$$

a   b   c

A. The reaction at equilibrium

b

$$k_c = \frac{[AB]^2}{[A_2][B_2]}$$

B. $Q_c < K_c$

C

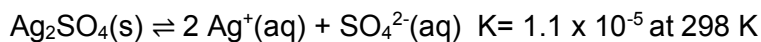$0.44 < 1.5$

C. $Q_c > K_c$

a

$36 > 1.5$

7

9. You are a given a solution of silver sulfate and the ionization of silver sulfate in water can be represented by the equation below

$$Ag_2SO_4(s) \rightleftharpoons 2\,Ag^+(aq) + SO_4^{2-}(aq) \quad K= 1.1 \times 10^{-5} \text{ at } 298\text{ K}$$

If the solution a 1.5 L solution contains 6.55 g Ag_2SO_4, and you attempt to dissolve additional solid silver sulfate in the solution will it dissolve?

A. Yes
B. No

B

molar mass of $Ag_2SO_4 = 311.799$ g/mol

$$K = [Ag^+]^2 [SO_4^{2-}]$$

* In order to determined if additional $AgSO_4$ will dissolve, you will have to determine the conc of Ag^+ ions & SO_4^{2-} then substitute the conc in the equilibrium expression (to determine Q & compare to K)

* if the solution is at equilibrium determined by K then the solution is saturated & no additional $AgSO_4$ will dissolve
   (Q = K)

* if Q<K ∴ additional Ag_2SO_4 will dissolve

Conc of Ag^+ & SO_4^{2-} ions

$$\frac{6.55\text{g }Ag_2SO_4}{1.5\text{L solution}} \times \frac{1\text{mol }Ag_2SO_4}{311.799\text{g }Ag_2SO_4} \times \frac{2\,Ag^+\text{ ions}}{Ag_2SO_4} = 0.0280094$$
$$\sim 0.280 \text{ M }Ag^+$$

$$\frac{6.55\text{g }Ag_2SO_4}{1.5\text{L solution}} \times \frac{1\text{mol }Ag_2SO_4}{311.799\text{g }Ag_2SO_4} \times \frac{SO_4^{2-}\text{ ions}}{Ag_2SO_4} = 0.0140047$$
$$\sim 0.0140 \text{ } SO_4^{2-} \text{ ions}$$

Substitute in $K = [0.280]^2 [0.0140]$
$$= 1.0987 \times 10^{-6}$$
$$\sim 1.1 \times 10^{-5} \sim \text{ the solution is at equilibrium } \therefore$$
saturated ∴
no more Ag_2SO_4 will dissolve

10. Phosgene, a colorless toxic gas ($COCl_2$) used in the synthesis of dyes and resins can be synthesized using carbon monoxide and chorine gases represented by the equilibrium reaction below

$$CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g) \quad K_c = 1.2 \times 10^3 \text{ at } 395\,°C$$
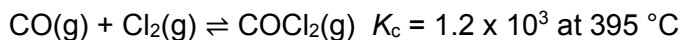
If 4.20 g CO(g) and 10.6 g Cl_2(g) are injected in a 3.25 L flask, what is the equilibrium concentration of the species below. Use **scientific notation** to report your answer.

A. Cl_2(g)

5.72 x 10^{-3} M

① Calculate the initial concentrations of CO & Cl_2

Molarity of CO = $4.20g\,CO \times \dfrac{1\,mol\,CO}{28.01g\,CO}$ $\dfrac{}{3.25L}$ = 0.04613736852

Molarity of Cl_2 = $10.6g\,Cl_2 \times \dfrac{1\,mol\,Cl_2}{70.90g\,Cl_2}$ $\dfrac{}{3.25L}$ = 0.04599546854

Calculating Q_c

= $\dfrac{[COCl_2]}{[CO][Cl_2]}$

= $\dfrac{[0]}{[0.0461][0.0460]}$ = 0

$Q_c < K_C$ ∴ reaction proceeds to the right

② Set up an ICE table to calculate the equilibrium concentrations

→ numbers are cut off due to limited space

I	0.0461	0.0460	0
C	$-x$	$-x$	$+x$
E	0.0461-x	0.0460-x	x

B. CO(g)

5.86 x 10^{-3} M

$[CO]_{eq} =$
$0.04613736852 - 0.0402728289$
$= 0.00586454462\,M$

③ Plug the equilibrium values from the ICE table into the equilibrium constant expression

$K_C = \dfrac{[COCl_2]}{[CO][Cl_2]}$

$1.2 \times 10^3 = \dfrac{x}{(0.0461-x)(0.0460-x)}$

$1.2 \times 10^3 (2.1221097 \times 10^{-3} - 9.213283406 \times 10^{-2} + x^2) = x$

$2.546531693 - 1.105594009 \times 10^2 x + 1.2 \times 10^3 x^2 = x$

$1.2 \times 10^3 x^2 - 1.116594009 \times 10^2 x + 2.546531693 = 0$

$x = 0.0402728289 \quad \text{or} \quad x = \cancel{0.0526933\,4352}$

C. $COCl_2$(g)

4.03 x 10^{-2} M

$[Cl_2]_{eq} = 0.04599546854 - 0.0402728289 = 0.0057226416\,M$

* Keep all your digits & round at the end to 3 sig figs

to check if your equilibrium concentrations are correct plug the equilibrium values you calculated back into the equilibrium constant expression

$k_c = \dfrac{[COCl_2]}{[CO][Cl_2]} = \dfrac{[0.0403]}{[0.00586][0.00572]}$
$= 1.20 \times 10^3$
↓
matches kc from the question

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11. A mixture of 0.250 mol $H_2(g)$ and 0.250 mol $I_2(g)$ in a 4.05 L flask is represented by the reaction below

$$H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g) \quad K_c = 50.2$$
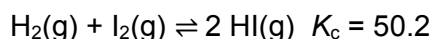
Calculate the equilibrium concentrations of the species below. Use **scientific notation** to report your answer.

A. $H_2(g)$

$\boxed{1.36}$ x 10$^{\boxed{-2}}$ M

① Conc. of H_2 & I_2

$$[H_2] = [I_2] = \frac{0.250\,mol}{4.05\,L} = 0.0617283950 6\,M$$

$Q = 0$ ∴ reaction proceeds towards products

numbers are truncated due to limited space

	H_2	I_2	
I	0.0617	0.0617	0
C	$-x$	$-x$	$+2x$
E	$0.0617-x$	$0.0617-x$	$2x$

② Plug the equilibrium conc. in k_c & solve for x

$$k_c = \frac{[HI]^2}{[H_2][I_2]}$$

$$50.2 = \frac{(2x)^2}{(0.0617-x)(0.0617-x)}$$

$$\sqrt{50.2} = \sqrt{\frac{4x^2}{(0.0617-x)^2}} \quad \text{taking the square root of both sides}$$

B. $I_2(g)$

$\boxed{1.36}$ x 10$^{\boxed{-2}}$ M

$$7.085195834 = \frac{2x}{0.0617-x}$$

$$2x = 7.0851958\,(0.0617-x)$$
$$2x = 0.4373577675 - 7.085195834\,x$$
$$9.085195834\,x = 0.4373577675$$
$$\therefore x = 0.0481396 0816$$

C. $HI(g)$

$\boxed{9.63}$ x 10$^{\boxed{-2}}$ M

③ $[H_2] = [I_2] = 0.0617283950 6 - 0.0481396 0816$
$$= 0.0135887869$$
$$= 0.0136\,M$$

$[HI] = 2x = 2(0.0481396 0816)$
$$= 0.0963\,M$$

check your answer:
$$k_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{[0.0963]^2}{[0.0136][0.0136]} = 50.1$$
↓ close to 50.2

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12. Which of the following statements is true regarding the equilibrium reaction $A(s) \rightleftharpoons B(s) + 2\ C(g) + \frac{1}{2} D(g)\ \Delta H = 0$?

E

A. The addition of excess solid $A(s)$ at equilibrium will not have any effect on the equilibrium mixture. ✓

B. If the volume of the equilibrium mixture is decreased by doubling the pressure, the system shifts to the left to reestablish equilibrium. ✓

C. Increasing or decreasing the temperature will not affect the equilibrium constant, K_c, for this reaction. ✓

D. If $Ne(g)$ is introduced to the equilibrium reaction at constant volume will .have no effect on the equilibrium position.

(E.) All of the statements above are true.

(A) Additions of solids & liquids at equilibrium has no effect on the equilibrium reaction

(B) Using the ideal gas law $PV = nRT$, $P \propto \frac{1}{V}$ & $P \propto n$
if $V \downarrow$: $P \uparrow$ & $\uparrow n$
∴ the system shifts to the left (fewer moles) to reestablish equilibrium

(C) $\Delta H = 0$ ∴ temperature has no effect on the equilibrium constant

(D) Introducing an inert gas at constant volume increases the total pressure but has no effect on the concentrations or partial pressures of the reactants or products

13. The decomposition of N_2O_4 to NO_2 is represented by the equilibrium reaction equation below

$$N_2O_4(g) \rightleftharpoons 2\ NO_2\ (g)$$
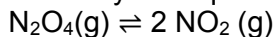
Initial concentration   *change*

If 1.75 moles of N_2O_4 is injected in a 1.5 L vessel and at equilibrium 25% N_2O_4 is dissociated, what is the equilibrium constant for this reaction?

0.389

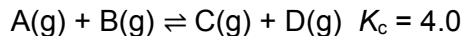|	I	1.17	0
	C	-0.292	+2(0.292)
	E	0.875	0.583

- Change in conc. of $N_2O_4 = \frac{25}{100} \times 1.17 = 0.29167$

- $K = \dfrac{[NO_2]^2}{[N_2O_4]} = \dfrac{[0.583]^2}{[0.875]}$

     $= 0.389$

or

If 25% of N_2O_4 is dissociated at equilibrium ∴ 75% of the initial conc. is left at equilibrium ∴

$\frac{75}{100} \times 1.17 = 0.875 \rightarrow$ you can plug this value in the ICE chart in the Equilibrium row & calculate the change for N_2O_4 & NO_2

14. For the hypothetical equilibrium reaction

$$A(g) + B(g) \rightleftharpoons C(g) + D(g) \quad K_c = 4.0$$

If you start with 0.373 M A(g), 0.396 M B(g), 0.552 M C(g), and 3.95 M D(g), what is the equilibrium concentration of C(g) and D(g)?

A. C(g)

| 0.357 | M

① Determine Q_c and compare to k_c to determine the direction of reaction

$$Q_c = \frac{[C][D]}{[A][B]} = \frac{[0.552][3.95]}{[0.373][0.396]} = 14.8$$

$Q_c > k_c$ ∴ the reaction proceeds towards reactants

② Set up an ICE table

I	0.373	0.396	0.552	3.95
C	$+x$	$+x$	$-x$	$-x$
E	$0.373+x$	$0.396+x$	$0.552-x$	$3.95-x$

③ $k_c = \dfrac{[C][D]}{[A][B]}$

$$4.0 = \frac{[0.552-x][3.95-x]}{[0.373+x][0.396+x]}$$

$$4.0 = \frac{2.18 - 4.50x + x^2}{0.148 + 0.769x + x^2}$$

B. D (g)

| 3.76 | M

$$4.0(0.148 + 0.769x + x^2) = 2.18 - 4.50x + x^2$$
$$0.592 + 3.08x + 4.0x^2 = 2.18 - 4.50x + x^2$$
$$3.0x^2 + 7.58x - 1.59 = 0$$
$$x = 0.195 \quad \text{or} \quad x = -2.72$$

$$[A] = 0.373 + 0.195 = 0.568$$
$$[B] = 0.396 + 0.195 = 0.591$$

Check your answer
$$k_c = \frac{[C][D]}{[A][B]}$$
$$= \frac{[0.357][3.76]}{[0.568][0.591]} = 3.998 \sim 4.0$$

$$[C] = 0.552 - 0.195 = 0.357 \, M$$
$$[D] = 3.95 - 0.195 = 3.76 \, M$$

15. If lead metal is added to a 0.100 M Cr^{3+}(aq) solution. What are the concentrations of Pb^{2+}(aq), Cr^{2+}(aq), and Cr^{3+}(aq) when the reaction is at equilibrium? Use **scientific notation** to report your answer ~~constant expression~~

↳ solids are not included in the equilibrium constant expression

$$Pb(s) + 2\ Cr^{3+}(aq) \rightleftharpoons Pb^{2+}(aq) + 2\ Cr^{2+}(aq) \quad K_c = 3.2 \times 10^{-10}$$
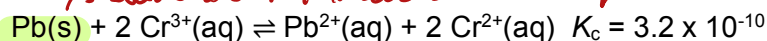
A. Pb^{2+}(aq)

$\boxed{9.28}$ x 10$\boxed{-5}$ M

I	–	0.100	0	0
C	–	– 2x	+x	+2x
E	–	0.100 – 2x	x	2x

① $K_C = \dfrac{[Pb^{2+}][Cr^{2+}]^2}{[Cr^{3+}]^2}$

$3.2 \times 10^{-10} = \dfrac{(x)(2x)^2}{(0.100 - 2x)^2}$

$\dfrac{C}{K} = \dfrac{0.100}{3.2 \times 10^{-10}} >> 100$ ∴ x is dropped in the denominator

② $3.2 \times 10^{-10} = \dfrac{(x)(2x)^2}{(0.100)^2}$

B. Cr^{2+}(aq)

$\boxed{1.86}$ x 10$\boxed{-4}$ M

$(0.100)^2 \times 3.2 \times 10^{-10} = 4x^2 \times x$

$3.2 \times 10^{-12} = 4x^3$

∴ $x^3 = 8.0 \times 10^{-13}$ → take the cube root of both sides

∴ $x = 9.283177667 \times 10^{-5}$

C. Cr^{3+}(aq)

$\boxed{9.98}$ x 10$\boxed{-2}$ M

$[Pb^{2+}] = x = 9.28 \times 10^{-5}$ M

$[Cr^{2+}] = 2x = 2(9.28 \times 10^{-5}) = 1.86 \times 10^{-4}$ M

$[Cr^{3+}] = 0.100 - 2(9.28 \times 10^{-5}) = 9.98 \times 10^{-2}$ M

13

16. The equilibrium mixture at 1000 K for the reaction $CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g)$ contains 0.276 M $H_2(g)$, 0.276 M $CO_2(g)$, 0.224 M $CO(g)$, and 0.224 M $H_2O(g)$. Determine the following:

A. The equilibrium constant, K_c

$\boxed{0.659}$

$$K_p = \frac{(PCO)(PH_2O)}{(PCO_2)(PH_2)}$$
$$= \frac{(0.224)(0.224)}{(0.276)(0.276)} = 0.659$$

B. $\Delta G°_{rxn}$ at 1000 K

$\boxed{3.49}$ kJ/mol

$$\Delta G_{rxn} = -RT \ln K_p$$
$$= -8.814 \frac{J}{mol \cdot K} \times 1000 \times \ln 0.657$$
$$= 3.49 \times 10^3 \frac{J}{mol}$$
$$= 3.49 \frac{kJ}{mol}$$

C. Does the reaction proceed towards products or reactants at 1000 K if a mixture contains 0.0750 M $CO_2(g)$, 0.095 M $H_2(g)$, 0.0340 mol $CO(g)$, and 0.0650 mol $H_2O(g)$?

$\boxed{I}$

I. The reaction proceeds towards products

II. The reaction proceeds towards reactants

III. The reaction is at equilibrium

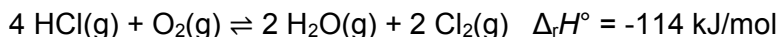$$Q_p = \frac{(PCO)(PH_2O)}{(PCO_2)(PH_2O)} = \frac{(0.0340)(0.0650)}{(0.0750)(0.0950)}$$
$$= 0.310$$
$$Q_p < K_p \quad \therefore \text{ the reaction proceeds to the right (toward products)}$$

14

17. One of the processes used in the production of chlorine gas is the Deacon process given by the equation below

$$4\ HCl(g) + O_2(g) \rightleftharpoons 2\ H_2O(g) + 2\ Cl_2(g) \quad \Delta_r H° = -114\ kJ/mol$$

If the above mixture is brought to equilibrium at 400 °C, which of the changes below will cause the system to increase the amount of $Cl_2(g)$? Select all that apply. Insert letters without spaces in the answer box, example **ABCD**.

ABD

A. ⬤ Cooling the reaction below 400 °C ✓

B. ⬤ Additional moles of HCl gas added to the system ✓

C. Increasing the volume of the system ✓

D. ⬤ Increasing the pressure of the system

E. Addition of a catalyst

Ⓐ ↓ in reaction temperature cause the reaction to shift to the right to produce more heat ∴ ↑ Cl_2 production

Ⓑ Addition of HCl ↓Q (↓Qc= $\frac{[H_2O]^2[Cl_2]}{[HCl]^4[O_2]}$), Qc < Kc, therefore the reaction shifts to the right ↑[HCl]⁴[O₂] producing more Cl_2

Ⓒ PV= nRT   ↓P=$\frac{n}{V}$RT → ↑ volume, ↓ pressure, ↓ no. of moles ∴ reaction shifts to node the side with more moles (left side ∴ ↓Cl_2) ↓ no. of moles

Ⓓ ↑ pressure ↑ no. of moles ∴ reaction shifts to the side with fewer moles (right side ∴ ↑ moles of Cl_2).

18. For the reaction $N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$, $\Delta H° = +57.2\ kJ/mol$ and $K = 0.113$ at 25 °C, what is the value of

A. K at 0 °C (report your answer using **scientific notation**)

$$1.36 \times 10^{-2}$$

Using the Van't Hoff Equation:

$$\ln\left(\frac{k_2}{k_1}\right) = \frac{\Delta H°}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$k_1 = ?$ $T_1 = 0°C + 273.15 = 273K$
$k_2 = 0.113$ $T_2 = 25 + 273.15 = 298\ k$
$\Delta H° = 57.2\ \frac{kJ}{mol} \times \frac{1000J}{1kJ} = 5.72 \times 10^4\frac{J}{mol}$

$$\ln\left(\frac{0.113}{k_1}\right) = \frac{5.72 \times 10^4\frac{J}{mol}}{8.314\frac{J}{mol \cdot k}}\left(\frac{1}{273} - \frac{1}{298}\right)$$

heat ←
↓ temp of reaction, reaction shifts to the left
↓ $K = \frac{[products]}{[reactants]}↑}$

$$\ln\left(\frac{0.113}{k_1}\right) = 2.114205052 \rightarrow \text{taking the inverse "e" of both sides}$$

$$\frac{0.113}{k_1} = e^{2.114205052}$$

$$\frac{0.113}{k_1} = 8.282998595$$

∴ $k_1 = 0.013642$
~ 1.36×10^{-2}

B. At what temperature will $K = 1.00$?

$$329\ K$$

Using the Van't Hoff Equation:

$$\ln\left(\frac{k_2}{k_1}\right) = \frac{\Delta H°}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$k_1 = 1.00, T_1 = ?$
$k_2 = 0.113$ $T_2 = 298$
$\Delta H = 5.72 \times 10^4\frac{J}{mol}$
$R = 8.314\frac{J}{mol \cdot k}$

$$\ln\left(\frac{0.113}{1.00}\right) = \frac{5.72 \times 10^4\frac{J}{mol}}{8.314\frac{J}{mol \cdot k}}\left(\frac{1}{T_1} - \frac{1}{298}\right)$$

$$-2.180367146 = 6.8799615 \times 10^3\left(\frac{1}{T_1} - \frac{1}{298}\right)$$

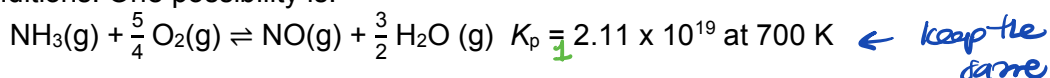$$-3.16915648 \times 10^{-4} = \frac{1}{T_1} - \frac{1}{298}$$

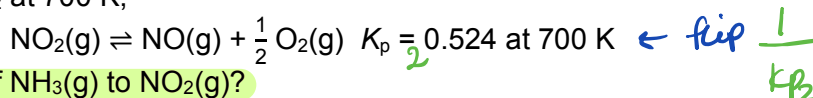$$\frac{1}{T_1} = 3.038783906 \times 10^{-3}$$

∴ $T_1 = 329.078$
~ 329 K

19. In the Ostwald process for oxidizing ammonia, a variety of products is possible – N_2, N_2O, NO, and NO_2 – depending on the conditions. One possibility is:
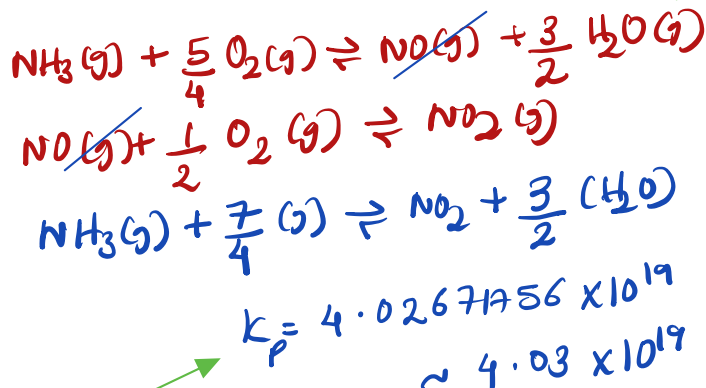
$$NH_3(g) + \frac{5}{4} O_2(g) \rightleftharpoons NO(g) + \frac{3}{2} H_2O \text{ (g) } K_p = 2.11 \times 10^{19} \text{ at 700 K}$$ ← keep the same

For the decomposition of NO_2 at 700 K,

$$NO_2(g) \rightleftharpoons NO(g) + \frac{1}{2} O_2(g) \text{ } K_p = 0.524 \text{ at 700 K}$$ ← flip $\frac{1}{kp_2}$

What is K_p for the oxidation of $NH_3(g)$ to $NO_2(g)$?

4.03 x 10 19

$$NH_3(g) + \frac{5}{4} O_2(g) \rightleftharpoons NO(g) + \frac{3}{2} H_2O(g)$$

$$NO(g) + \frac{1}{2} O_2(g) \rightleftharpoons NO_2(g)$$

$$NH_3(g) + \frac{7}{4}(g) \rightleftharpoons NO_2 + \frac{3}{2}(H_2O)$$

$$K_p = 4.0267456 \times 10^{19}$$
$$\sim 4.03 \times 10^{19}$$

$$\cdot k_p = K_{p_1} \times \frac{1}{K_{p_2}}$$

20. The equilibrium reaction for the synthesis of 1,2 dichloroethane is given below

$$C_2H_4(g) + Cl_2(g) \rightleftharpoons C_2H_4Cl_2(g) \text{ } \Delta H = -1244 \text{ kJ/mol}$$

2 mol (g)   1 mol (g)

To maximize the amount of $C_2H_4Cl_2(g)$ produced, which of the following strategies might be applied? Assume the reaction is at equilibrium. Select all that apply. Insert letters without spaces in the answer box, example **ABCD.**

BCO

A. Increasing the reaction volume
B. Removing $C_2H_4Cl_2(g)$ from the reaction mixture as it forms ✓
C. Adding $Cl_2(g)$ ✓
D. Lowering the reaction temperature ✓
E. None of the choice above will increase the amount of $C_2H_4Cl_2(g)$ produced

Ⓐ ↑ in volume ↓ Pressure ↓ no. of moles ∴ the reaction proceeds towards the side with greater number of moles (to the left) ∴ ↓ $C_2H_4Cl_2$ production

Ⓑ Removing $C_2H_4Cl(g)$ from the reaction ↓ its conc ∴ the reaction proceeds to the right to produce more C_2H_4Cl

Ⓒ ↑ $Cl_2(g)$ would cause the reaction to proceed towards the right ↑ $C_2H_4Cl_2(g)$ production

Ⓞ ↓ in reaction temperature cause the reaction to shift to the right to produce more heat ∴ ↑ $C_2H_4Cl_2$ production

16

21. The free energy change of the reaction A(g) → B(g) is zero under certain conditions. The standard free energy of the reaction is -42.5 kJ. Which statement is true about the reaction? Select all that apply. Insert $\Delta G^\circ =$ letters without spaces in the answer box, example **ABCD**.

$\Delta G_{rxn}^\circ < 0$ (spontaneous) ∴ k > 1

-42.5 kJ/mol

A. The concentration of the product is greater than the concentration of the reactant.

B. The reaction is at equilibrium. $\Delta G_{rxn} = 0$

C. The concentration of the reactant is greater than the concentration of the product. inunet
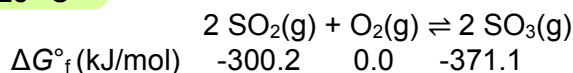
D. None of the above statements is correct

Answer box: **AB**

$\Delta G = 0$ same as A

$k = \dfrac{[Product]^n}{[reactant]^n}$ T

$\Delta G = \Delta G_{rxn}^\circ + RT \ln Q$

At equilibrium ← $\Delta G_{rxn}^\circ + RT \ln Q = 0$

∴ $\Delta G_{rxn}^\circ = -RT \ln Q$

22. You are given the information below for the equilibrium reaction between sulfur dioxide and oxygen to produce sulfur trioxide at 25 °C

$$2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$$

	2 SO₂(g)	O₂(g)	2 SO₃(g)
ΔG°_f (kJ/mol)	-300.2	0.0	-371.1

If $SO_2(g)$ has a partial pressure is 1.0×10^{-4} atm, $O_2(g)$ is 0.20 atm, and $SO_3(g)$ is 0.10 atm, calculate ΔG_{rxn}.

Answer box: **-104** kJ/mol

① Calculate ΔG_{rxn}°

$\Delta G_{rxn}^\circ = \Sigma n \Delta G_f\, products - \Sigma n \Delta G_f\, reactants$

$= (2 \times -371.1) - [(2 \times -300.2) + 0.0]$

$= -141.8$ kJ/mol

② Calculate Q_p

$Q_p = \dfrac{(P_{SO_3})^2}{(P_{SO_2})^2 (O_2)} = \dfrac{(0.10)^2}{(1.0 \times 10^{-4})^2 (0.20)}$

$= 5.0 \times 10^6$

③ $\Delta G_{rxn} = \Delta G_{rxn}^\circ + RT \ln Q$

$= \left(-141.8 \dfrac{kJ}{mol}\right) + \left[\left(8.314 \dfrac{J}{mol \cdot K} \times \dfrac{1 kJ}{1000 J}\right) \times 298 \times \ln 5.0 \times 10^6\right)\right]$

$= -104$ kJ/mol

Extra Practice Questions: these questions will not be graded

1. You are given two elementary reactions below which are the reverse of one another and the reaction rates for elementary reaction are provided as well. Which of the following statements about an equilibrium mixture of X, Y, and Z is **false**?

$$X + Y \rightarrow Z + Z \quad \text{Rate} = k_1 [X][Y]$$
$$Z + Z \rightarrow X + Y \quad \text{Rate} = k_2 [Z]^2$$

At equilibrium
Rate of forward = Rate of reverse
$$k_1 [X][Y] = k_2 [Z]^2$$

A. At equilibrium, k_1 and k_2 can have different values
B. At equilibrium, $k_1 [X][Y] = k_2 [Z]^2$
C. At equilibrium, X and Y still react to form Z
D. At equilibrium, the concentrations of X and Y must be the same
E. At equilibrium, $\frac{[X][Y]}{[Z]^2}$ is a constant
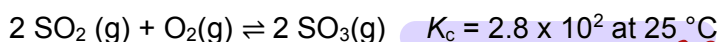
D (answer boxed)

k_1 & k_2 can have different values but when multiplied by their respective conc, the rates will be equal to each other

the concentration of reactants is not equal at equilibrium

2. For the reaction,

$$2 SO_2 (g) + O_2(g) \rightleftharpoons 2 SO_3(g) \quad K_c = 2.8 \times 10^2 \text{ at } 25\,°C$$

Do not forget to calculate the molarity of reactants & products

If a mixture contains 0.455 mol of SO_2, 0.183 mol of O_2, and 0.568 mol SO_3, are introduced into a 1.90 L vessel at 727 K. Which of the following statements is true?

A. The reaction is at equilibrium because $Q_c = K_c$
B. The reaction will proceed towards the reactants because $Q_c < K_c$
C. The reaction will proceed towards the products because $Q_c > K_c$
D. Not enough information is provided in the question

C (answer boxed)

to determine if the reaction is at equilibrium you will need to calculate Q_c & compare it to K_c
- if $Q_c = K_c$ ∴ the reaction is at equilibrium
- if $Q_c < K_c$ ∴ the reaction proceeds towards the products (to the right)
- if $Q_c > K_c$ ∴ the reaction proceeds towards the reactants (to the left)
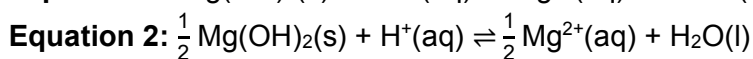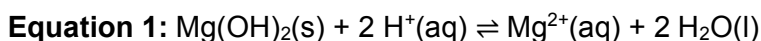
$$[SO_2] = \frac{0.455}{1.90} = 0.23947868$$
$$[O_2] = \frac{0.183}{1.90} = 0.096315789$$
$$[SO_3] = \frac{0.568}{1.90} = 0.298947868$$

$$Q_c = \frac{[SO_3]^2}{[SO_2]^2 [O_2]} = \frac{[0.298947868]^2}{[0.23947868]^2[0.096315789]}$$
$$= 16.17992072$$

$Q_c < K_c$ ∴ reaction proceeds towards products

18

3. Two equations can be written for the dissolution of $Mg(OH)_2(s)$ in acidic solution

Equation 1: $Mg(OH)_2(s) + 2 H^+(aq) \rightleftharpoons Mg^{2+}(aq) + 2 H_2O(l)$

Equation 2: $\frac{1}{2} Mg(OH)_2(s) + H^+(aq) \rightleftharpoons \frac{1}{2} Mg^{2+}(aq) + H_2O(l)$
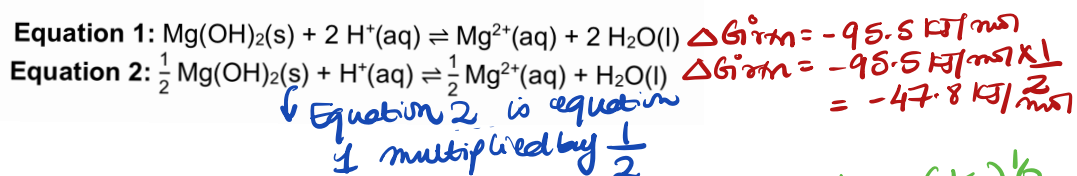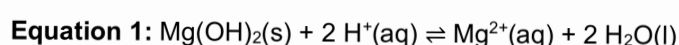
Which of the statements below <mark>is/are correct regarding the relationship between the two equations?</mark> <mark>Select all that apply.</mark> Insert letters without spaces in the answer box, example **ABCD**.
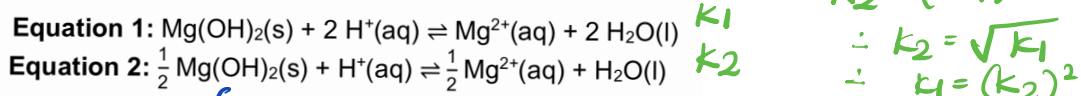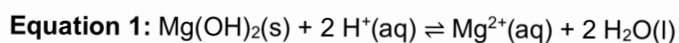
AC

A. If $\Delta G°_{rxn}$ for equation 1 is −95.5 kJ/mol therefore $\Delta G°_{rxn}$ for equation 2 will be −47.8 kJ/mol

B. If $\Delta G°_{rxn}$ for equation 1 is −95.5 kJ/mol therefore $\Delta G°_{rxn}$ for equation 2 will be 9120 kJ/mol

C. K_1, the equilibrium constant for equation 1 is the square of K_2, the equilibrium constant for equation 2

D. K_1, the equilibrium constant for equation 1 is the square root of K_2, the equilibrium constant for equation 2

E. None of the above statements is/are correct

Equation 1: $Mg(OH)_2(s) + 2 H^+(aq) \rightleftharpoons Mg^{2+}(aq) + 2 H_2O(l)$ $\Delta G°_{rxn} = -95.5 \text{ kJ/mol}$

Equation 2: $\frac{1}{2} Mg(OH)_2(s) + H^+(aq) \rightleftharpoons \frac{1}{2} Mg^{2+}(aq) + H_2O(l)$ $\Delta G°_{rxn} = -95.5 \text{ kJ/mol} \times \frac{1}{2}$
$= -47.8 \text{ kJ/} \frac{2}{mol}$

✓ Equation 2 is equation 1 multiplied by $\frac{1}{2}$

Equation 1: $Mg(OH)_2(s) + 2 H^+(aq) \rightleftharpoons Mg^{2+}(aq) + 2 H_2O(l)$ K_1

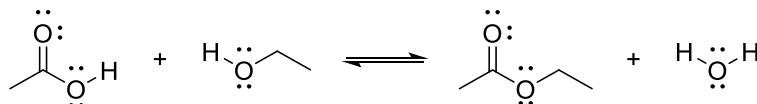**Equation 2:** $\frac{1}{2} Mg(OH)_2(s) + H^+(aq) \rightleftharpoons \frac{1}{2} Mg^{2+}(aq) + H_2O(l)$ K_2

$K_2 = (K_1)^{\frac{1}{2}}$
∴ $K_2 = \sqrt{K_1}$
∴ $K_1 = (K_2)^2$

4. The following shows an example of an organic chemistry equilibrium reaction known as the Fisher Esterification.

E

No effect on products and reactants

Which of the following would correctly <mark>predict the effect of adding an acid catalyst to this reaction?</mark>

A. The reaction equilibrium would change from favoring the products to favoring the reactants

B. The reaction equilibrium would change from favoring the reactants to favoring the products

C. The reaction equilibrium would have a higher concentration of products than without the catalyst

D. The reaction equilibrium would have a higher concentration of reactants than without the catalyst

E. The reaction would reach a state of equilibrium faster but the equilibrium concentrations would not change

F. There is not enough information to determine how a catalyst would affect the reaction

* As a reminder, once the reaction reaches equilibrium, the catalyst has NO effect on the rate of the equilibrium reaction as the rate of the forward & reverse processes are the same.

* A catalyst will help the reaction only reach the equilibrium state faster

5. The Citric acid cycle is a metabolic pathway that connects carbohydrate, fat, and protein metabolism. One important reaction in the Citric acid cycle is:

Citrate(aq) $\rightleftharpoons$ aconitate (aq) + $H_2O(l)$ *+ heat* $K = 3.1 \times 10^2$ $\Delta H = +6.7$ kJ/mol
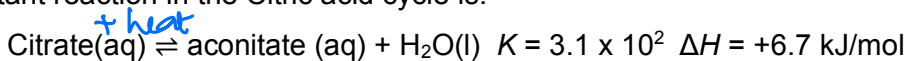
Which of the statements regarding the equilibrium reaction is **true**? Select all that apply.

ACD

A. The equilibrium reaction favors the formation of aconitate. *$K > 1$ ∴ reaction is product favored*

B. If the enzyme catalyst, aconitase is added the equilibrium concentration of aconitate increases. *Catalysts have no effect on the speed or concentration of the reaction*

C. If the reaction is heated at 40 °C, the concentration of aconitate increases.

D. Adding more $H_2O(l)$ has no effect on the equilibrium position.

C) reaction is endothermic ∴ increasing the temperature of the reaction by heating it to 40°C shifts the equilibrium to the right increasing the concentration of aconitate

D) Adding solids & liquids have no effect on the equilibrium position

6. Which of the following statements are **false** regarding the following reaction, given the equilibrium constant shown?

$$A(g) \rightleftharpoons B(g) \quad K_c = 10 \quad \text{Product favored}$$
$$K > 1$$

D

A. At equilibrium, the reaction is product favored
B. If [A] = 1.0 M and [B] = 1.0 M, then the reaction is not at equilibrium; the concentration of [B] will increase as the reaction moves toward equilibrium
C. If [A] = 0.1 M and [B] = 1.0 M, then the reaction is already at equilibrium, and the concentrations of products and reactants will not change
D. If [A] = 1.0 M and [B] = 1.0 M, then the reaction is not at equilibrium; the concentration of [A] will increase as the reaction moves toward equilibrium
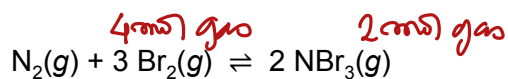
A) true

B) $Q = \dfrac{[B]}{[A]} = \dfrac{[1.0]}{[1.0]} = 1.0$ $Q < K$ ∴ reaction proceed toward B, product

C) $Q = \dfrac{[1.0]}{[0.1]} = 10$ ∴ $Q = K$ ∴ the reaction is at equilibrium be at equilibrium the conc of reactants & products do not change

D) see explanation of B

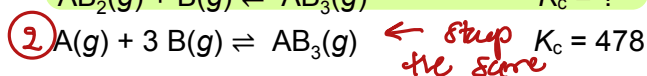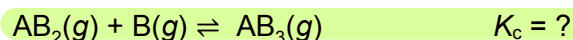7. What is Δn for the following equation in relating K_c to K_p?

$N_2(g) + 3\ Br_2(g) \rightleftharpoons 2\ NBr_3(g)$

(handwritten: 4 mol gas above $N_2 + 3Br_2$, 2 mol gas above $2NBr_3$)
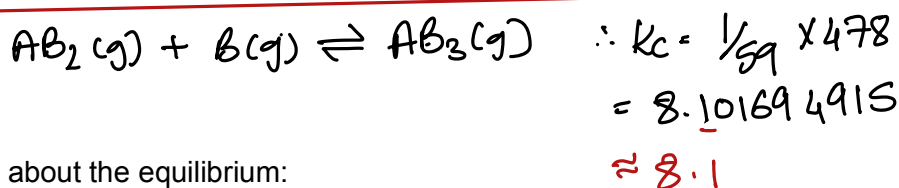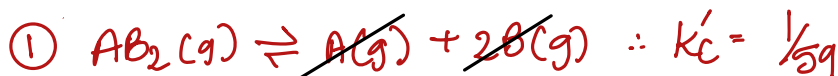
(handwritten: $\Delta n = 2 - 4 = -2$)

C

A. 4
B. 3
C. -2
D. 1
E. 0

8. The equilibrium constant is given for two of the reactions below. Determine the value of the missing equilibrium constant.
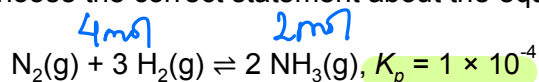
C

① $A(g) + 2\ B(g) \rightleftharpoons AB_2(g)$ ← flip $K_c = 59$ ∴ $\frac{1}{K_C} = \frac{1}{59}$

$AB_2(g) + B(g) \rightleftharpoons AB_3(g)$ $K_c = ?$

② $A(g) + 3\ B(g) \rightleftharpoons AB_3(g)$ ← keep the same $K_c = 478$

A. 3.5×10^{-5}
B. 2.8×10^4
C. 8.1
D. 0.12
E. 89

① $AB_2(g) \rightleftharpoons A(g) + 2B(g)$ ∴ $K_c' = \frac{1}{59}$

② $A(g) + 3B(g) \rightleftharpoons AB_3(g)$ ∴ $K_c = 478$

$AB_2(g) + B(g) \rightleftharpoons AB_3(g)$ ∴ $K_c = \frac{1}{59} \times 478$

$= 8.10169 4915$

≈ 8.1

9. Choose the correct statement about the equilibrium:

(handwritten: 4 mol, 2 mol)

$N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g),\ K_p = 1 \times 10^{-4}$

D

A. The rate constant for the forward reaction is greater than that of the reverse reaction. **X**
B. Since the reaction has a high activation energy, a catalyst is not needed.
C. The equilibrium constant is given by $K_p = [N_2][H_2]^3/[NH_3]^2$.
D. Conducting the reaction under high pressures will increase the yield of ammonia.
E. The K_p is independent of temperature.

A) $K = \frac{k_f}{k_r}$ *(rate constant)* if $K < 1$ ∴ $k_f < k_r$

(G, equilibrium constant)

B) K gives no information on the kinetics of the equilibrium process

C) $K_p = (PNH_3)^2 / [N_2][H_2]^3$

D) $PV = nRT$, $P \propto n$ ∴ ↑P, ↑n & the equilibrium reaction shifts to the side with fewer moles ∴ producing more NH_3 moles

E) K_p is temperature dependent

21

10. Choose the INCORRECT statement.

[**B**]

A. The van't Hoff equation is $\ln \frac{K2}{K1} = \frac{\Delta_r H°}{R}\left[\frac{1}{T1} - \frac{1}{T2}\right]$. *True*

B. K_{eq} is independent of temperature. *False*

C. In a thermodynamic equilibrium constant expression, the activity of a gas is replaced by its partial pressure in atmosphere.

D. In a K_{eq} expression, the activity of an aqueous species can be approximated by the numerical value of its molarity. *True*

E. If $\Delta G = 0$, the process is at equilibrium. *True*

$$\Delta G_{rxn} = \Delta G°_{rxn} + RT \ln K$$
at equilibrium $\Delta G_{rxn} = 0$
$$\therefore 0 = \Delta G°_{rxn} + RT \ln K$$
$$\therefore \Delta G°_{rxn} = -RT \ln K$$

11. For the reaction: $CO(g) + 2 H_2(g) \rightarrow CH_3OH(g)$

K_p = 91.4 at 350 K and K_p = 2.05 × 10^{-4} at 298 K.

What is the value of $\Delta H°_{rxn}$?

[**D**]

A. 49.9 kJ/mol
B. 2.08 × 10^3 kJ/mol
C. 3.74 × 10^{-2} kJ/mol
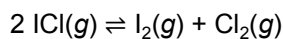D. 217 kJ/mol
E. 446 kJ/mol

~ Van't Hoff equation

$$\ln \frac{K_2}{K_1} = \frac{\Delta H°_{rxn}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\ln\left(\frac{2.05 \times 10^{-4}}{91.4}\right) = \frac{\Delta H°_{rxn}}{8.314 \times 10^{-3} \text{ J/mol·K}}\left(\frac{1}{350} - \frac{1}{298}\right)$$

$$\therefore \Delta H°_{rxn} = 216.916723 \text{ kJ/mol}$$
$$\sim 217 \text{ kJ/mol}$$
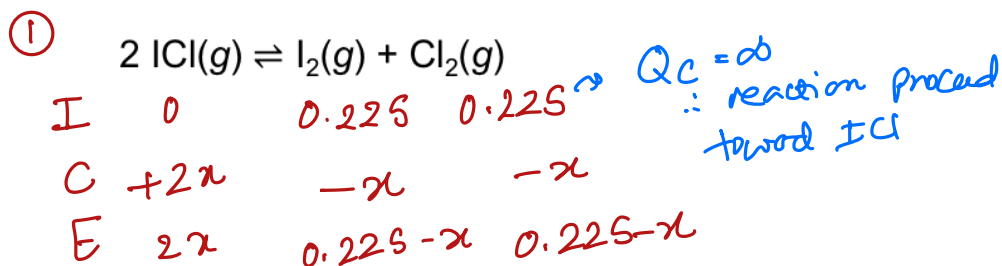
12. At a certain temperature the equilibrium constant, K_C, equals 0.11 for the reaction:

C

$$2 ICl(g) \rightleftharpoons I_2(g) + Cl_2(g)$$

What is the equilibrium concentration of ICl if 0.45 mol of I_2 and 0.45 mol of Cl_2 are initially mixed in a 2.0 L flask?

A. 0.14 mol/L
B. 0.17 mol/L
C. 0.27 mol/L
D. 0.34 mol/L

③ $0.331662479 =$

$\dfrac{0.225 - x}{2x}$

∴ $0.6633299581x$
$= 0.225 - x$
∴ $1.6633299581 x =$
0.225
∴ $x = 0.1352712$

④ $[ICl] = 2x = 0.27 \, mol/L$

① $2 ICl(g) \rightleftharpoons I_2(g) + Cl_2(g)$

	I		
I	0	0.225	0.225
C	$+2x$	$-x$	$-x$
E	$2x$	$0.225-x$	$0.225-x$

$Q_c = \infty$
∴ reaction proceed
toward ICl

② $K_c = \dfrac{[I_2][Cl_2]}{[ICl]^2}$

∴ $0.11 = \dfrac{[0.225-x][0.225-x]}{[2x]^2}$

∴ $\sqrt{0.11} = \sqrt{\dfrac{[0.225-x]^2}{[2x]^2}}$

Formula Sheet

Length
1 kilometer = 0.62137 mile
1 inch = 2.54 centimeters (exactly)
1 Ångstrom = 1×10^{-10} meter

Energy
1 joule = $1 \text{ kg·m}^2/\text{s}^2$
1 calorie = 4.184 joules
1 Calorie = 1 kilocalorie = 1000 calories
1 L·atm = 101.325 joules

Pressure
1 pascal = $1 \text{ N/m}^3 = 1 \text{ kg/m·s}^2$
1 atmosphere = 101.325 kilopascals = 760 mm Hg = 760 torr = 14.70 lb/in^2
1 bar = 1×10^5 Pa (exactly)

Temperature
0 K = −273.15°C
K = °C + 273.15
°C = (5/9)(°F − 32)

Mass
1 kg = 2.205 lbs

Volume
1 mL = 1 cm^3 = 1 cc

Constants
$c = 2.998 \times 10^8$ m/sec
$h = 6.626 \times 10^{-34}$ J·sec^{-1}
R = 0.08206 L·atm/mol·K = 8.314 J/mol·K
Specific heat of water = 4.184 J/g·K
Mass of an electron: 9.109×10^{-31} kg
Mass of a proton: 1.673×10^{-27} kg
RH = 2.18×10^{-18} J
Specific heat of water = 4.184 J/g·K
STP = 273.15 K and 1 atm
Avogadro's number: 6.022×10^{23}

Equations
d (density) = m/V
$P_1V_1 = P_2V_2$
$V_1/T_1 = V_2/T_2$
$P_1V_1/n_1T_1 = P_2V_2/n_2T_2$
PV = nRT
$(P + a(n^2/V^2))·(V - nb) = nRT$
molar mass (M) = mRT/PV
density (d) = MP/RT
$x_A = n_A/n_{tot} = P_A/P_{tot} = V_A/V_{tot}$
$P_{tot} = P_A + P_B + ...$
$n_{tot} = n_A + n_B + ...$

$$\mu_{rms} = \sqrt{\frac{3RT}{M}}$$

$$\frac{Rate\ of\ effusion\ A}{Rate\ of\ effusion\ B} = \sqrt{\frac{MW_B}{MW_A}}$$

$Q = C \times \Delta T = c_{specific} \times m \times \Delta T$

$Q = n \times \Delta H$ (kJ/mol) $= m \times \Delta H$ (kJ/g)

$w = -P\Delta V$

$\Delta E = q + w$

$\Delta H° = \Sigma n\Delta Hf°(products) - \Sigma n\Delta Hf°(reactants)$

$\Delta H° = \Sigma n\Delta H°(bonds\ broken) - \Sigma n\Delta H°(bonds\ formed)$

$E = h\nu$

$c = \lambda\nu$

$\lambda = h/m\nu$

$$\Delta E = -2.18 \times 10^{-18} J\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{vap}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$C_g = kP_g$

$P_{solution} = P_{solvent} X_{solvent}$

$P_{solution} = \Sigma P_j = \Sigma P_j X_j$

$\Delta T_b = K_b m i$

$\Delta T_f = K_f m i$

$\pi = MRT i$

Thermodynamic and Electrochemistry

$S = k_b \times \ln(W)$

$k_b = 1.381 \times 10^{-23}$ J/K

$\Delta S = q_{rev}/T$

$\Delta S_{surr} = q_{surr}/T = -q_{rev}/T$

$\Delta S_{univ} = \Delta S_{sys} + \Delta S_{surr}$

$\Delta S°_{rxn} = \Sigma v S°_{products} - \Sigma v S°_{reactants}$

$\Delta H°_{rxn} = \Sigma v H°_{products} - \Sigma v H°_{reactants}$

$\Delta G°_{rxn} = \Sigma v G°_{products} - \Sigma v G°_{reactants}$

$\Delta G = \Delta H - T\Delta S$

$\Delta G = \Delta G° + RT \cdot \ln Q$

R=8.314 J/mol.K

$\Delta G° = -RT \cdot \ln K$

$\Delta G = -nFE_{cell}$

$F = 96485$ J/(V·mol e$^-$)

$E°_{cell} = RT/nF \ln K$

$E°_{cell} = (0.0257/n) \ln K = (0.0592/n) \log K$

$E_{cell} = E°_{cell} - (RT/nF) \ln Q$

$E_{cell} = E°_{cell} - (0.0257/n) \ln Q$

Electrolysis: Q (total charge) = I × t = n × F

Integrated Rate Laws & half-life

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

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

$$[A] = -kt + [A]_0$$

$$t_{1/2} = \frac{[A]_0}{2k}$$

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

$$t_{1/2} = \frac{1}{k[A]_0}$$

$$\ln \frac{k_2}{k_1} = -\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

Equilibrium and Acid / Base

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

$pH = -\log[H_3O^+]$

$K_w = 1.0 \times 10^{-14}$ at 25 °C

$K_w = [H_3O^+] \times [OH^-]$

$K_w = K_a \times K_b$

$pK_a = -\log[K_a]$

Buffer: $pH = pK_a + \log \frac{[A^-]}{[HA]}$

$$\ln \frac{K_2}{K_1} = \frac{\Delta H^\circ_{rxn}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

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** 1.01																	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** 24.31	3	4	5	6	7	8	9	10	11	12	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** 47.87	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** 65.38	31 **Ga** 69.72	32 **Ge** 72.63	33 **As** 74.92	34 **Se** 78.97	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** 95.95	43 **Tc** [97]	44 **Ru** 101.07	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** 121.76	52 **Te** 127.60	53 **I** 126.90	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** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	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]	89 **Ac** [227]	104 **Rf** [267]	105 **Db** [268]	106 **Sg** [269]	107 **Bh** [270]	108 **Hs** [269]	109 **Mt** [277]	110 **Ds** [281]	111 **Rg** [282]	112 **Cn** [285]	113 **Nh** [286]	114 **Fl** [290]	115 **Mc** [290]	116 **Lv** [293]	117 **Ts** [294]	118 **Og** [294]

57 **La** 138.91	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** [145]	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97
89 **Ac** [227]	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]