

Recitation Worksheet Ten: Exam 3 Review

Name:

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Textbook:

Chemistry & Chemical Reactivity

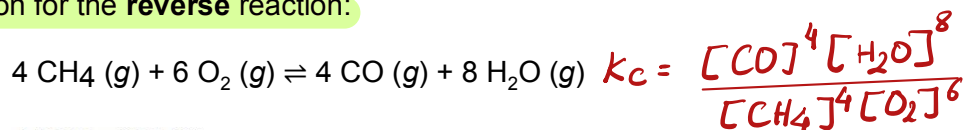
by John C. Kotz, Paul M. Treichel, John R. Townsend, David Treichel

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Instructions:

- This recitation worksheet is a review for Exam One.
- Exam coverage: 15.1-15.6, 18.5-18.7, and 16.1-16.3
- You **do not** need to submit it to Gradescope.
- The answer key has been posted with this worksheet to eLC.
- The **recitation session during the exam week (October 28th– 31st) is still mandatory**. Your attendance will be recorded.
- A periodic table and formula sheet are attached to the end of this worksheet.

1. Write the equilibrium equation for the **reverse** reaction:



0

A. $K_c' = \frac{4 [\text{CO}] + 8 [\text{H}_2\text{O}]}{4 [\text{CH}_4] + 6 [\text{O}_2]}$

B. $K_c' = \frac{4 [\text{CO}_4] + 6 [\text{O}_2]}{4 [\text{CO}] + 8 [\text{H}_2\text{O}]}$

C. $K_c' = \frac{[\text{CO}]^4 [\text{H}_2\text{O}]^8}{[\text{CH}_4]^4 [\text{O}_2]^6}$

D. $K_c' = \frac{[\text{CH}_4]^4 [\text{O}_2]^6}{[\text{CO}]^4 [\text{H}_2\text{O}]^8}$

the reverse reaction is:

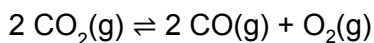


K_c for this reaction

$$= \frac{[\text{CH}_4]^4 [\text{O}_2]^6}{[\text{CO}]^4 [\text{H}_2\text{O}]^8}$$

2. Determine the value of K_p for the following reaction if the equilibrium partial pressures are as follows:
 $P(\text{CO}_2)_{\text{eq}} = 1.8 \text{ atm}$, $P(\text{CO})_{\text{eq}} = 0.35 \text{ atm}$, $P(\text{O}_2)_{\text{eq}} = 0.50 \text{ atm}$. (Do not use scientific notation for your answer).

0.019



$$K_p = \frac{P(\text{CO})^2 P(\text{O}_2)}{P(\text{CO}_2)^2}$$

$$= \frac{(0.35)^2 (0.50)}{(1.8)^2}$$

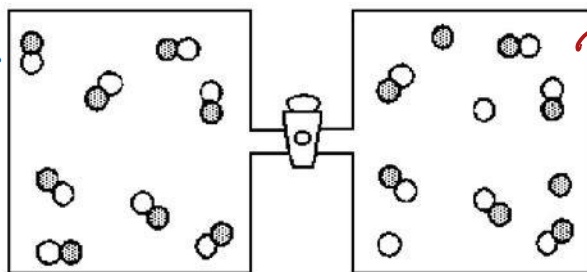
$$= 0.01890432099$$

$$\sim 0.019$$

3. Consider the reaction $A + B \rightleftharpoons 2 AB$. The vessel on the right contains an equilibrium mixture of A atoms (shaded spheres), B atoms (unshaded spheres), and AB molecules.

If the barrier is removed & the content of the vessel on the left is allowed to mix with the equilibrium mixture

\therefore you need to calculate Q_c & compare it to K_c



\rightarrow Equilibrium

$$K_c = \frac{[AB]^2}{[A][B]}$$

$$= \frac{[6]^2}{[2][2]} = \frac{36}{4} = 9$$

If the barrier between the two vessels is removed and the contents of the two vessels are allowed to mix, what will be observed?

0

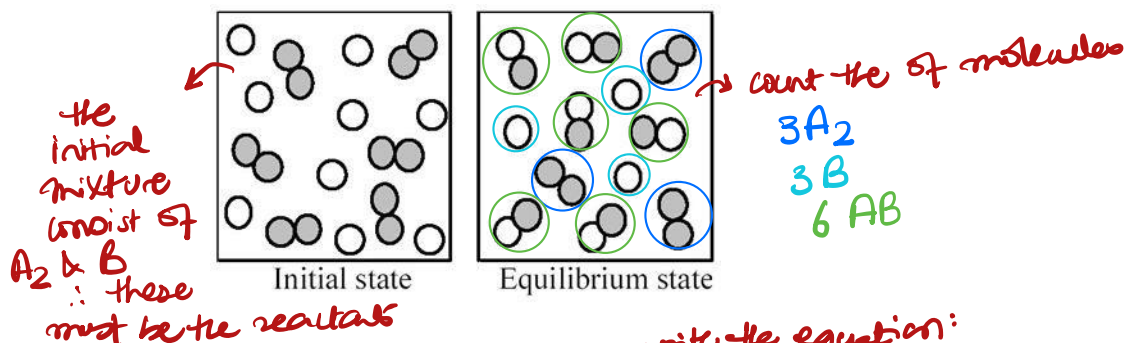
- A. The reaction will go in the forward direction decreasing the number of A atoms and B atoms and increasing the number of AB molecules.
- B. The reaction will go in the forward direction increasing the number of A atoms and B atoms and decreasing the number of AB molecules.
- C. The reaction will go in the reverse direction decreasing the number of A atoms and B atoms and increasing the number of AB molecules.
- ☒ D. The reaction will go in the reverse direction increasing the number of A atoms and B atoms and decreasing the number of AB molecules.

\rightarrow From the vessel to the left

$$Q_c = \frac{[AB]^2}{[A][B]} = \frac{[8+6]^2}{[2][2]} = 49$$

$\therefore Q_c > K_c$ & the reaction shifts to the left (reactants)
 $[AB] \downarrow$ & $[A] \uparrow$ & $[B] \uparrow$

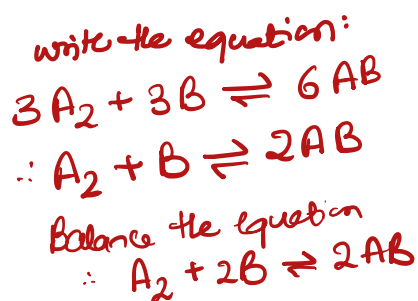
4. The following pictures represent the initial state and the equilibrium state for the gaseous state reaction of A_2 molecules (shaded spheres) with B atoms (unshaded spheres) to give AB molecules.



What is the best balanced chemical equation for the reaction?

C

- A. $A_2 + B \rightleftharpoons A_2B$
 B. $A_2 + 2 B \rightleftharpoons A_2B_2$
 C. $A_2 + 2 B \rightleftharpoons 2 AB$
 D. $6 A_2 + 9 B \rightleftharpoons 3 A_2 + 3 B + 6 AB$



5. At a certain temperature, bromine and nitric oxide react to form nitrosyl bromide:
 $Br_2(g) + 2 NO(g) \rightleftharpoons 2 NOBr(g)$

When 0.010 mol Br_2 is mixed with 0.025 mol NO and 0.015 mol $NOBr$ in a 2.50 L flask, the concentration of $NOBr$ decreases. Which statement below is true?

C

- A. $K_C < 36$
 B. $K_C > 36$
 C. $K_C < 90$
 D. $K_C > 90$

$$Q_C = \frac{[NOBr]^2}{[Br_2][NO]^2}$$

$$= \frac{\left[\frac{0.015}{2.50}\right]^2}{\left[\frac{0.010}{2.50}\right]\left[\frac{0.025}{2.50}\right]^2} = \frac{3.6 \times 10^{-5}}{4.0 \times 10^{-7}} = 9.0 \times 10^1$$

if the conc. of $NOBr$ decreases at equilibrium
 $\therefore K_C < 90$

6. Given:



Which of the following would be true if the temperature were increased from 25 °C to 100 °C?

1. The value of K_c would be smaller.
2. The concentration of A(g) would be decreased.
3. The concentration of B(g) would increase.

↳ Altering the temperature of the reaction changes the value of K_c

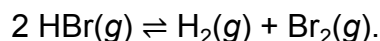
E

- A. 1 only
- B. 2 only
- C. 3 only
- D. 1 and 2 only
- E. 2 and 3 only

** If the temperature of an endothermic reaction is raised, the reaction shifts toward the right (products)*
** If the reaction shifts toward products: concentration of the products increases & that of the reactants decreases & the value of K_c increases*

$$K_c = \frac{[B][C]}{[A]^2} \uparrow$$

7. Gaseous hydrogen bromide decomposes at elevated temperatures according to the equation:



At a certain temperature a 2.00 L flask is initially filled only with 0.600 mol of HBr. What is the value of K_c at that temperature if the flask contains 0.104 mol of H_2 at equilibrium?

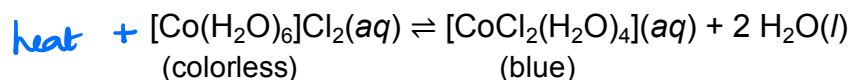
A

- A. 7.04×10^{-2}
- B. 4.40×10^{-2}
- C. 3.00×10^{-2}
- D. 2.10×10^{-1}

	$[\text{HBr}]$	$[\text{H}_2]$	$[\text{Br}_2]$
Initial	0.300	0	0
Change	-2x	+0.0520	+x
Equilibrium	0.196	0.0520	0.0520

$$\begin{aligned}
 K_c &= \frac{[\text{H}_2][\text{Br}_2]}{[\text{HBr}]^2} \\
 &= \frac{[0.0520][0.0520]}{[0.196]^2} \\
 &= \frac{0.002704}{0.038416} = 0.0703873386 \\
 &\approx 7.04 \times 10^{-2}
 \end{aligned}$$

8. A crude type of disappearing ink is based on the following **endothermic equilibrium**:



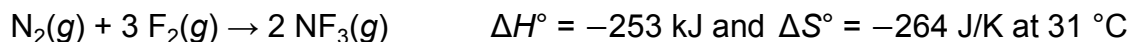
If the reactant solution is used to write on a piece of paper and the paper is allowed to partially dry, what can be done to bring out the colored handwriting?

D

- A. add water
B. decrease the volume
C. put the paper in a freezer
D. put the paper in an oven

to bring out the colored handwriting, the reaction shifts to the right.
If the piece of paper is heated (in an oven), the reaction will shift to the right to dissipate the heat \therefore the blue color of the handwriting will show.

9. Consider the reaction:



Calculate ΔG° and state whether the equilibrium composition should favor reactants or products at standard conditions.

C

- A. $\Delta G^\circ = -333 \text{ kJ}$; the equilibrium composition should favor products.
 B. $\Delta G^\circ = -333 \text{ kJ}$; the equilibrium composition should favor reactants.
 C. $\Delta G^\circ = -173 \text{ kJ}$; the equilibrium composition should favor products.
 D. $\Delta G^\circ = -173 \text{ kJ}$; the equilibrium composition should favor reactants.

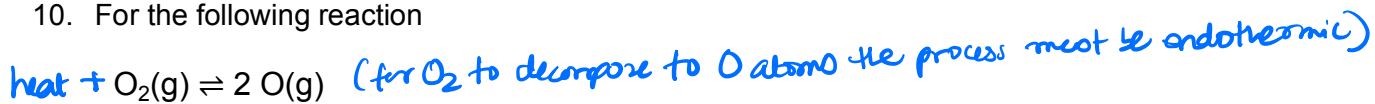
$$\begin{aligned}\Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ \\ &= -253 \text{ kJ} - (31 + 273 \text{ K})(-264 \times 10^{-3} \text{ kJ}) \\ &= -172.744 \sim -173 \text{ kJ}\end{aligned}$$

stants.

Also $\Delta G^\circ = -RT \ln K_P$
if $\Delta G^\circ < 0 \therefore K > 1$
 $\Delta G^\circ > 0 \quad K < 1$
 $\Delta G^\circ = 0 \quad K \approx 1$

If $\Delta G^\circ < 0$ \therefore the reaction is spontaneous (product favored) \therefore the equilibrium composition should favor products

10. For the following reaction



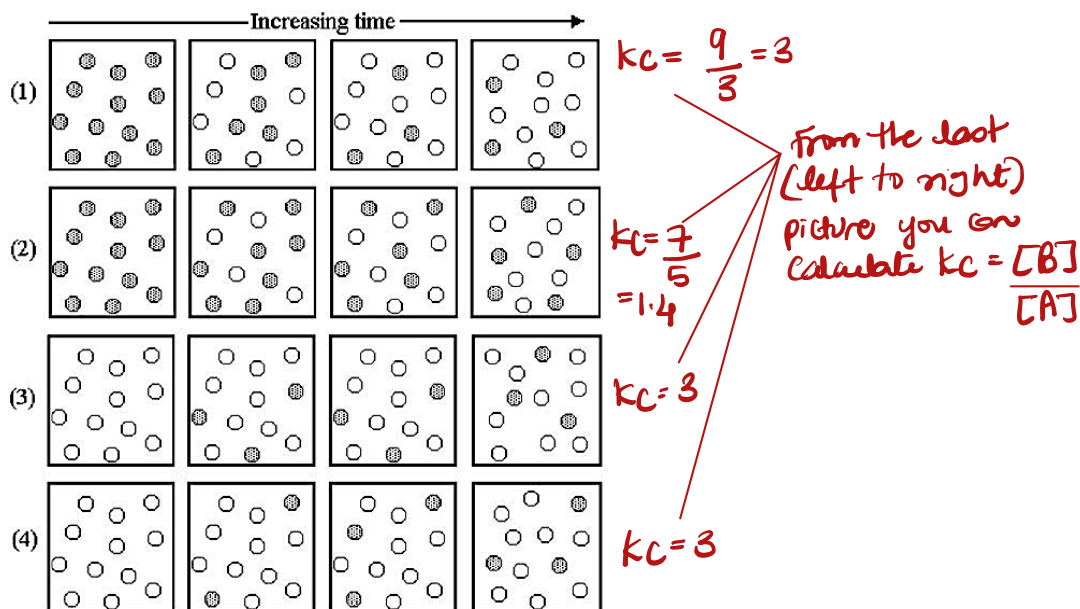
what conditions favor production of oxygen atoms?

A

- A. high temperature and low pressure
- B. high temperature and high pressure
- C. low temperature and low pressure
- D. low temperature and high pressure

- If the temperature of an endothermic reaction is raised \therefore the reaction shifts to the right to dissipate the heat increasing the production of oxygen atoms → high temperature
- If the pressure of the vessel increases $\&$ P \downarrow n \therefore if the pressure of the vessel decreases, the number of mol of gas in the system decrease $\&$ the reaction shifts to the side with more moles of gas (to the right) to reestablish equilibrium \therefore the production of oxygen atoms is favored → low pressure

11. Consider the interconversion of A molecules (shaded spheres) and B molecules (unshaded spheres) according to the reaction $\text{A} \rightleftharpoons \text{B}$. Each of the series of pictures represents a separate experiment in which time increases from left to right.

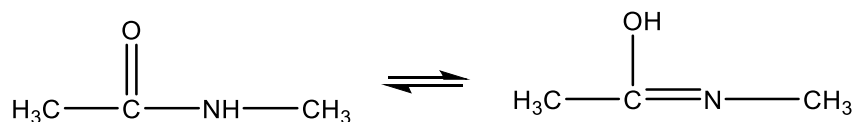


Which of these experiments has resulted in an equilibrium state?

B

- A. all of the experiments except experiment (1)
- B. all of the experiments except experiment (2)
- C. all of the experiments except experiment (3)
- D. all of the experiments except experiment (4)

12. In solution, the two C_3H_7NO isomers exist in equilibrium:



If $K_c = 0.57$ at $31^\circ C$ and the initial concentration of the reactant is $0.50 M$ and the product is $0.70 M$, what are the concentrations at equilibrium?

C

- A. [reactant] = $0.43 M$ and [product] = $0.24 M$
 B. [reactant] = $0.67 M$ and [product] = $0.38 M$
 C. [reactant] = $0.76 M$ and [product] = $0.44 M$
 D. [reactant] = $0.82 M$ and [product] = $0.47 M$

	[reactant]	[product]
Initial	0.50	0.70
Change	+x	-x
Equilibrium	0.50+x	0.70-x

$$\begin{aligned} \therefore 0.57 [0.50 + x] &= [0.70 - x] \\ 0.285 + 0.57x &= 0.70 - x \\ \therefore 1.57x &= 0.415 \\ \therefore x &= 0.26 \end{aligned}$$

$$\therefore [\text{reactant}] = 0.50 + 0.26 = 0.76 M \text{ \& } [\text{product}] = 0.70 - 0.26 = 0.44 M$$

$$\begin{aligned} 1) \text{ Calculate } Q_c &= \frac{[\text{Product}]}{[\text{Reactant}]} \\ &= \frac{0.70}{0.50} = 1.4 \end{aligned}$$

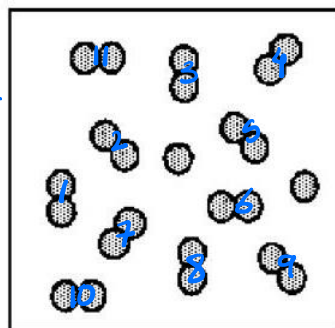
$\therefore Q_c > K_c$ & the reaction shifts to the left (towards reactants)

$$\begin{aligned} 2) \text{ Calculate } x \text{ using } K_c \\ K_c &= \frac{[\text{Product}]}{[\text{Reactant}]} \\ 0.57 &= \frac{[0.70 - x]}{[0.50 + x]} \end{aligned}$$

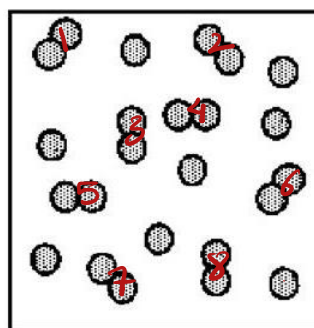
13. Consider the reaction $2A(g) \rightleftharpoons A_2(g)$. The pictures represent the initial state and the equilibrium state of the system.

For initial state calculate Q_p

$$\begin{aligned} \therefore Q_p &= \frac{P(A_2)}{P(A)^2} \\ &= \frac{(11)}{(2)^2} = \frac{11}{4} \\ &= 2.75 \end{aligned}$$



Initial State



Equilibrium State

For equilibrium state calculate K_p

$$\begin{aligned} \therefore K_p &= \frac{P(A_2)}{P(A)^2} \\ &= \frac{(8)}{(8)} = 1 \end{aligned}$$

For initial state 2 what is the relationship between the reaction quotient, Q_p , and the equilibrium constant, K_p ?

D

- A. $Q_p < K_p$
 B. $Q_p = K_p = 1$
 C. $Q_p = K_p \neq 1$
 D. $Q_p > K_p$

$$Q_p > K_p$$

14. At 2600 K, $\Delta G^\circ = 775 \text{ kJ}$ for the vaporization of boron carbide: $\text{B}_4\text{C}(\text{s}) \rightleftharpoons 4 \text{B}(\text{g}) + \text{C}(\text{s})$. Find ΔG and determine if the process is spontaneous if the reaction vessel contains 4.00 mol $\text{B}_4\text{C}(\text{s})$, 0.400 mol of $\text{C}(\text{s})$, and $\text{B}(\text{g})$ at a partial pressure of $1.00 \times 10^{-5} \text{ atm}$. At this temperature, $R \times T = 21.6 \text{ kJ}$. Keep 3 significant figures.

-220. kJ

$$\Delta G = \Delta G^\circ + RT \ln Q_p$$

\downarrow 775 kJ \nwarrow 21.6 kJ

$$Q_p = (1.00 \times 10^{-5})^4 = 1.00 \times 10^{-20}$$

$$\therefore \Delta G = 775 \text{ kJ} + (21.6 \text{ kJ} \times \ln(1.00 \times 10^{-20}))$$

$$= -219.7167602 \text{ kJ}$$

$$\approx -220. \text{ kJ}$$

3 sig figs

15. For a reaction at constant temperature, as Q increases

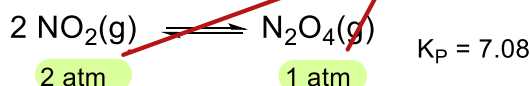
C

- A. ΔG and ΔG° increase.
 B. ΔG and ΔG° decrease.
 C. ΔG increases, but ΔG° remains constant.
 D. ΔG decreases, but ΔG° remains constant.

At a constant temperature there is only one ΔG° value
 $\therefore \Delta G^\circ$ (standard free energy) remains constant

As Q increases in comparison to K (remains constant at constant temperature)
 \therefore the process becomes reactant favored & ΔG increases

16. The following reaction is observed to have the provided partial pressures. Which condition best describes the state of the reaction.



B

- A. The reaction is at equilibrium
 B. The reaction is forming products faster than it is forming reactants
 C. The reaction is forming reactants faster than it is forming products
 D. The reaction is proceeding to the products rapidly
 E. The reaction is proceeding to the reactants rapidly

A) From the partial pressures you can calculate Q_p

$$Q_p = \frac{P(\text{N}_2\text{O}_4)}{P(\text{NO}_2)^2} = \frac{(1)}{(2)^2} = 0.25$$

$\therefore Q_p < K_p$
 \therefore the reaction is not at equilibrium

B) Looking at the equilibrium constant, $K_p = 7.08$, & if $K = \frac{k_f}{k_r}$
 $\therefore k_f > k_r$ & products are forming faster than reactants

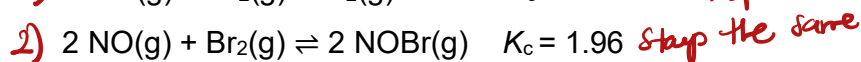
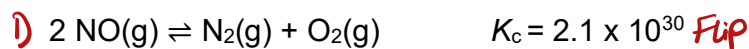
C) if $K_p > 1 \therefore k_f > k_r \therefore$ incorrect

D) if $Q_p < K_p$ it indicates spontaneity not the speed at which the reaction proceeds

$$K = \frac{k_{f \rightarrow \text{rate}} \text{ not at forward}}{k_{r \rightarrow \text{rate}} \text{ not at reverse process}}$$

$$K_p = 7.08 > 1$$

17. The value for the equilibrium constant, K_c , is given for the following reactions:



What is K_c for $\text{N}_2\text{(g)} + \text{O}_2\text{(g)} + \text{Br}_2\text{(g)} \rightleftharpoons 2 \text{NOBr(g)}$ at 298 K?

0

A. 9.7×10^{-16}

B. 2.1×10^{30}

C. 4.1×10^{30}

D. 9.3×10^{-31}

E. 8.7×10^{-61}



$\therefore K_c = 9.3333333 \times 10^{-31}$
 $\approx 9.3 \times 10^{-31}$

18. What is the value of K_p in question 17 for $\text{N}_2\text{(g)} + \text{O}_2\text{(g)} + \text{Br}_2\text{(g)} \rightleftharpoons 2 \text{NOBr(g)}$ at 298 K?

3.82×10^{-32}

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

$K_p = 9.3 \times 10^{-31} (0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 298)^{2-3}$

$\therefore K_p = 3.81670857 \times 10^{-32}$
 $\approx 3.82 \times 10^{-32}$

constant

constant

\therefore
 infinit sig
 figs

2 sig figs
 based
 on
 temperature

19. What is the value of ΔG° in kJ/mol for the reaction $2 \text{SO}_2\text{(g)} + \text{O}_2\text{(g)} \rightleftharpoons 2 \text{SO}_3\text{(g)}$ at 545 K given the following concentrations: $[\text{SO}_2]_{\text{eq}} = 0.0911 \text{ M}$, $[\text{O}_2]_{\text{eq}} = 0.0822 \text{ M}$, $[\text{SO}_3]_{\text{eq}} = 0.0982 \text{ M}$ at 545 K?

5.22

kJ

$\Delta G^\circ = -RT \ln K_p$ \hookrightarrow gaseous phase reaction

$R = 8.314 \times 10^{-3} \frac{\text{kJ}}{\text{mol} \cdot \text{K}}$
 $T = 545 \text{ K}$
 $K_p = ?$

1) to calculate K_p you need to calculate K_c & then convert to K_p

$\therefore K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]} = \frac{[0.0982]^2}{[0.0911]^2 [0.0822]} = 14.1356051$

$\therefore K_p = K_c (RT)^{\Delta n} = 14.1356051 (0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 545)^{2-3}$
 $= 0.3160722653$

2) calculate ΔG°

$\Delta G^\circ = -8.314 \times 10^{-3} \frac{\text{kJ}}{\text{mol}} \times 545 \times \ln(0.3160722653)$
 $= 5.218884866 \text{ kJ/mol} \approx 5.22$

2 sig figs based on temperature

20. At equilibrium, a 1.0-liter container was found to contain 0.20 moles of A, 0.20 moles of B, 0.40 moles of C and 0.40 mole of D. If 0.10 moles of A and 0.10 moles of B are added to this system, what will be the new equilibrium concentration of A? *Equilibrium is disturbed ∴ you have to recalculate the initial conc*

0.23

M
2)

	[A]	[B]	[C]	[D]
Initial	0.30	0.30	0.40	0.40
Change	-x	-x	+x	+x
Equilibrium	0.30-x	0.30-x	0.40+x	0.40+x

$$A(g) + B(g) \rightleftharpoons C(g) + D(g)$$

1) Calculate $K_c = \frac{[C][D]}{[A][B]} = \frac{[0.40][0.40]}{[0.20][0.20]} = \frac{0.16}{0.040} = 4$

Calculate Q_c to compare it to K_c to determine the direction of the reaction

$$Q_c = \frac{[C][D]}{[A][B]} = \frac{[0.40][0.40]}{[0.30][0.30]} = 1.8$$

$Q_c < K_c$ ∴ the reaction proceeds toward products

3) solve for x using K_c

$$4 = \frac{[0.40+x][0.40+x]}{[0.30-x][0.30-x]} \quad \therefore \sqrt{4} = \frac{[0.40+x]^2}{[0.30-x]^2}$$

$$\therefore 2 = \frac{0.40+x}{0.30-x} \quad \therefore 2(0.30-x) = 0.40+x$$

$$0.60 - 2x = 0.40 + x$$

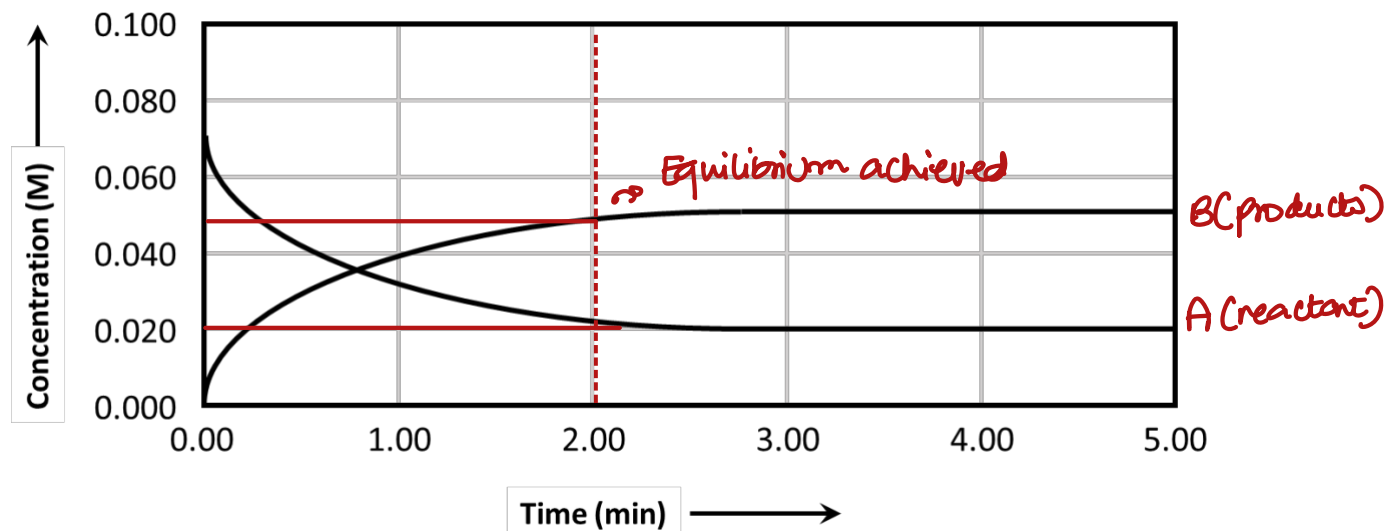
$$0.20 = 3x$$

$$\therefore x = 0.067$$

$$4) [A] = 0.30 - x = 0.30 - 0.067$$

$$= 0.23$$

21. Shown below is a concentration vs. time plot for the reaction $A \rightleftharpoons B$. For this reaction the value of the equilibrium constant is



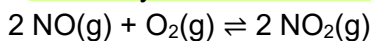
D

- A. $K_c < 1$.
B. $K_c = 0$.
C. $K_c = 1$.
D. $K_c > 1$.

$$K_c = \frac{[B]}{[A]} = \frac{[0.050]}{[0.020]} = 2.5$$

∴ $K_c > 1$

22. Consider this reaction. At equilibrium, 6.00 mol of NO and 1.90 mol of O₂ are present at equilibrium in a 2.50 L flask. If the value for K_c is 23.8, how many moles of NO₂ are present?



25.5

mol

$$1) [\text{NO}] = \frac{6.00 \text{ mol}}{2.50 \text{ L}} = 2.40 \text{ M}$$

$$[\text{O}_2] = \frac{1.90 \text{ mol}}{2.50 \text{ L}} = 0.760 \text{ M}$$

$$2) K_c = \frac{[\text{NO}_2]^2}{[\text{NO}]^2 [\text{O}_2]} \quad \therefore 23.8 = \frac{[\text{NO}_2]^2}{[2.40]^2 [0.760]^2}$$

$$\therefore \sqrt{[\text{NO}_2]^2} = \sqrt{104.18688}$$

$$\therefore [\text{NO}_2] = \frac{10.20719746 \text{ mol}}{2.50 \text{ L}} = 25.51799365 \approx 25.5 \text{ mol}$$

23. Consider the equilibrium system: N₂O₄(g) ⇌ 2 NO₂(g) for which K_p = 0.1134 at 25 °C and ΔH_r[°] = 58.03 kJ/mol. Assume that 1 mole of N₂O₄ and 2 moles of NO₂ are introduced into a 5.0-liter container. Calculate the equilibrium value of [N₂O₄]?

0.379

M

1)

	[N ₂ O ₄]	[NO ₂]
Initial	0.2	0.4
Change	+x	-2x
Equilibrium	0.2+x	0.4-2x

$$\rightarrow Q_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = \frac{[0.4]^2}{[0.2]} = 0.8$$

2) Calculate K_c from K_p

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

$$\therefore 0.1134 = K_c \left(\frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 298 \right)^{2-1}$$

$$\therefore K_c = 4.64 \times 10^{-3}$$

Q_c > K_c ∴ reaction proceeds in the reverse direction

3) Calculate x by plugging K_c

$$4.637300911 \times 10^{-3} = \frac{[0.4 - 2x]^2}{[0.2 + x]}$$

$$4.637300911 \times 10^{-3} [0.2 + x] = [0.4 - 2x]^2$$

$$9.274601822 \times 10^{-4} + 4.637300911 \times 10^{-3} x = 0.16 - 1.6x + 4x^2$$

$$4x^2 - 1.604637301x + 0.1590725398 = 0$$

$$\therefore x = 0.1790374689 \text{ or } x = 0.2221218663$$

$$\therefore [\text{N}_2\text{O}_4] = 0.2 + 0.179 = 0.379 \text{ M}$$

disregard this value

24. Consider the equilibria:

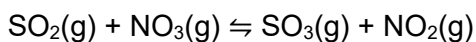


$$K_c = 2.3 \times 10^{-7}$$



$$K_c = 1.4 \times 10^{-3}$$

Calculate the equilibrium constant for the reaction



A

A. 78

B. 1.3×10^{-2}

C. 1.6×10^{-4}

D. 3.2×10^{-10}

E. 6.1×10^3

$$1) \text{SO}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightleftharpoons \text{SO}_3(\text{g}) \quad \therefore K_c = \left(\frac{1}{2.3 \times 10^{-7}} \right)^{\frac{1}{2}}$$

$$2) \text{NO}_3(\text{g}) \rightleftharpoons \text{NO}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \quad \therefore K_c = (1.4 \times 10^{-3})^{\frac{1}{2}}$$



$$\therefore K_c = \left(\frac{1}{2.3 \times 10^{-7}} \right)^{\frac{1}{2}} \times (1.4 \times 10^{-3})^{\frac{1}{2}}$$

$$= 78.01894976$$

$$\approx 78$$

25. The equilibrium constant for the reaction $\frac{1}{2} \text{Cl}_2(\text{g}) + \frac{1}{2} \text{F}_2(\text{g}) \rightleftharpoons \text{ClF}(\text{g})$ is measured to be 9.3×10^9 at 298 K and 3.3×10^7 at 398 K. Calculate:

A. $\Delta G^\circ_{\text{rxn}}$ at 298 K for the reaction in kJ/mol

-56.9

kJ/mol

$$\Delta G^\circ_{\text{rxn}} = -RT \ln K_p \quad \rightarrow \text{sig figs based on temperature}$$

$$= -8.314 \times 10^{-3} (298) \times \ln(9.3 \times 10^9)$$

$$= -56.86840442$$

$$\approx -56.9 \text{ kJ/mol}$$

B. ΔH° between 298 K and 398 K

-55.6

kJ/mol

\rightarrow sig figs are based on temperature

$$\ln \frac{K_2}{K_1} = \frac{\Delta H^\circ_{\text{rxn}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right) \quad \rightarrow \text{Van't Hoff equation}$$

$$K_1 = 9.3 \times 10^9$$

$$T_1 = 298$$

$$K_2 = 3.3 \times 10^7$$

$$T_2 = 398$$

$$R = 8.314 \times 10^{-3} \text{ kJ/mol} \cdot \text{K}$$

$$\Delta H^\circ = ?$$

$$\therefore \ln \left(\frac{3.3 \times 10^7}{9.3 \times 10^9} \right) = \frac{\Delta H^\circ_{\text{rxn}}}{8.314 \times 10^{-3}} \left(\frac{1}{298} - \frac{1}{398} \right)$$

$$\therefore \Delta H^\circ_{\text{rxn}} = -55.62699961 \text{ kJ/mol}$$

$$\approx -55.6 \text{ kJ/mol}$$

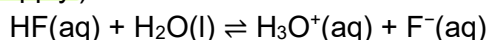
26. Which of these pairs is **NOT** a Brønsted-Lowry conjugate acid-base pair?

B

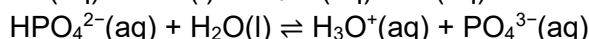
- A. HOBr and OBr⁻
 B. H₂SO₄ and SO₄²⁻ → difference is two protons
 C. C₆H₅NH₃⁺ and C₆H₅NH₂
 D. H₂CO₃ and HCO₃⁻
 E. C₆H₅CH₂CO₂H and C₆H₅CH₂CO₂⁻

the difference between a Brønsted-Lowry conjugate acid-base pair is only one proton (H⁺)

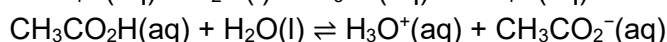
27. For these three acids and their acid equilibrium constants, which of the statements are true? (Select all that apply.)



$$K_a = 7.2 \times 10^{-4}$$



$$K_a = 3.6 \times 10^{-13}$$



$$K_a = 1.8 \times 10^{-5}$$

B O E

- A. HPO₄²⁻ is the strongest acid *False: has the smallest equilibrium constant*
 B. HF produces the weakest conjugate base out of the three acids
 C. The rate of dissociation of HF is the fastest as indicated by its acid equilibrium constant *False: K_a gives information about the extent of the dissociation of an acid*
 D. The conjugate base of CH₃CO₂H is CH₃CO₂⁻ *True*
 E. HF is the strongest acid *True (largest K_a)*
 B. *True: HF is the strongest acid out of the three acids ∴ produces the weakest conjugate base*
 D. *True: the difference between a conjugate acid-base pair is one proton*

28. You are given four solutions A, B, C, and D. Solution A has an [H₃O⁺] = 1.7 × 10⁻⁴ M, solution B has a pH = 5.25, solution C has an [OH⁻] = 1.5 × 10⁻⁵ and solution D has a pOH = 1.55. Arrange the solutions in the order of increasing [H₃O⁺].

B

- A. A < B < C < D
 B. D < C < B < A
 C. C < B < A < D
 D. B < A < D < C
 E. A < B < D < C

A) Solution A: [H₃O⁺] = 1.7 × 10⁻⁴ M

B) Solution B: pH = 5.25
 [H₃O⁺] = 10^{-pH} = 10^{-5.25} = 5.6 × 10⁻⁶ M

C) Solution C: [OH⁻] = 1.5 × 10⁻⁵
 [H₃O⁺][OH⁻] = 1.0 × 10⁻¹⁴
 ∴ [H₃O⁺][1.5 × 10⁻⁵] = 1.0 × 10⁻¹⁴
 ∴ [H₃O⁺] = 6.7 × 10⁻¹⁰

D) Solution D: pOH = 1.55
 ∴ pH + pOH = 14.00
 pH = 14.00 - pOH = 14.00 - 1.55
 = 12.45
 [H₃O⁺] = 10^{-12.45} = 3.5 × 10⁻¹³ M

Solutions in increasing order of [H₃O⁺]:
 D < C < B < A

smallest to largest

29. A solution of a weak acid is prepared by dissolving 0.040 mol of HA in sufficient water to yield 2.0-L of solution. The pH of the solution was 3.53 at 25.0 °C. Calculate the K_a of HA.

B

- A. 2.9×10^{-4}
 B. 4.4×10^{-6}
 C. 8.7×10^{-8}
 D. 1.7×10^{-11}
 E. 1.1×10^{-6}

i)



	[HA]	[H ₂ O]	[A ⁻]	[H ₃ O ⁺]
Initial	2.0×10^{-2}	—	0	0
Change	-3.0×10^{-4}	—	$+3.0 \times 10^{-4}$	$+3.0 \times 10^{-4}$
Equilibrium	2.0×10^{-2}	—	3.0×10^{-4}	3.0×10^{-4}

→ [H₃O⁺]_{eq} can be determined from pH

∴ [H₃O⁺]_{eq} = $10^{-\text{pH}}$ (2 decimal places)
 $= 10^{-3.53}$
 $= 3.0 \times 10^{-4}$
 (2 sig figs)

$$\begin{aligned} 2) K_a &= \frac{[A^-][H_3O^+]}{[HA]} \\ &= \frac{[3.0 \times 10^{-4}][3.0 \times 10^{-4}]}{[2.0 \times 10^{-2}]} \\ &= 4.4 \times 10^{-6} \end{aligned}$$

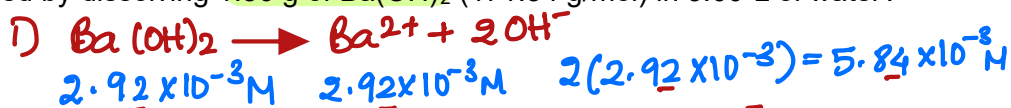
* Note: Do not round until the end

→ strong base

30. What is the pH of a solution prepared by dissolving 1.50 g of Ba(OH)₂ (171.34 g/mol) in 3.00-L of water?

B

- A. 2.23
 B. 11.77
 C. 11.47
 D. 2.53
 E. 13.70



2) $pOH = -\log [OH^-] = -\log [5.84 \times 10^{-3}] = 2.234$

3) $pH + pOH = 14.000$

∴ $pH = 14.000 - pOH = 14.000 - 2.234 = 11.766$ (3 decimal places)

Molarity of Ba(OH)₂:

$$\frac{1.50 \text{ g } Ba(OH)_2 \times \frac{1 \text{ mol } Ba(OH)_2}{171.34 \text{ g } Ba(OH)_2}}{3.00 \text{ L}}$$

$= 2.91817429 \times 10^{-3} M$ (3 sig figs)

31. Find the pH of a 0.135 M aqueous solution of a weak acid HA, for which $K_a = 2.3 \times 10^{-2}$.

0

- A. 1.25
B. 3.28
C. 1.17
D. 1.34
E. 1.64



	$[HA]$	$[H_2O]$	$[A^-]$	$[H_3O^+]$
Initial	0.135	—	0	0
Change	$-x$	—	$+x$	$+x$
Equilibrium	$0.135 - x$	—	x	x

2) $K_a = \frac{[A^-][H_3O^+]}{[HA]}$

$2.3 \times 10^{-2} = \frac{[x][x]}{[0.135 - x]}$

$2.3 \times 10^{-2} [0.135 - x] = x^2$

$3.105 \times 10^{-3} - 2.3 \times 10^{-2} x = x^2$

$\therefore x^2 + 2.3 \times 10^{-2} x - 3.105 \times 10^{-3} = 0$

$\therefore x = 0.04539688647$

or $x = -0.068896886 \rightarrow$

disregard this value

$\hookrightarrow \frac{C}{K} < 100 \therefore$ cannot drop x in the denominator

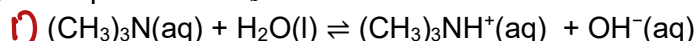
3) $[H_3O^+] =$

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

$= -\log [0.0454]$

$= 1.343$

32. An aqueous solution of a weak base, trimethylamine, $(CH_3)_3N$, has a pH = 10.50. What will be the concentration of $(CH_3)_3N$ at equilibrium? $K_b = 6.3 \times 10^{-5}$.



$1.6 \times 10^{-3} M$

	$[(CH_3)_3N]$	$[H_2O]$	$[(CH_3)_3NH^+]$	$[OH^-]$
Initial	x	—	0	0
Change	-3.2×10^{-4}	—	$+3.2 \times 10^{-4}$	$+3.2 \times 10^{-4}$
Equilibrium	$x - 3.2 \times 10^{-4}$	—	3.2×10^{-4}	3.2×10^{-4}

Can be determined from pH

if $pH = 10.50$

$\therefore pOH = 14.00 - 10.50 = 3.50$

$\therefore [OH^-]_{eq} = 10^{-pOH} = 10^{-3.50} = 3.16227766 \times 10^{-4}$

2) $K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]}$

$6.3 \times 10^{-5} = \frac{[3.2 \times 10^{-4}][3.2 \times 10^{-4}]}{[x - 3.2 \times 10^{-4}]}$

$6.3 \times 10^{-5} [x - 3.2 \times 10^{-4}] = [3.2 \times 10^{-4}]^2$

$\therefore x = 1.903529353 \times 10^{-3}$

$\therefore [(CH_3)_3N]_0 = 1.9 \times 10^{-3} = x$

16

$\therefore [CH_3N]_{eq} = 1.903529353 \times 10^{-3} - 3.16227766 \times 10^{-4} = 1.587301687 \times 10^{-3} \approx 1.6 \times 10^{-3} M$

33. What is the conjugate base of sulfuric acid?

A

- A. HSO_4^-
- B. H_3O^+
- C. OH^-
- D. SO_4^{2-}
- E. H_3SO_4^+

H_2SO_4 (acid) \therefore conjugate base is one proton less $\therefore \text{HSO}_4^-$

34. What is the $[\text{H}_3\text{O}^+]$ for a solution at 25°C that has $\text{pOH} = 5.640$?

C

- A. $2.34 \times 10^{-4} \text{ M}$
- B. $2.29 \times 10^{-6} \text{ M}$
- C. $4.37 \times 10^{-9} \text{ M}$
- D. $4.27 \times 10^{-11} \text{ M}$
- E. 8.360 M

$\text{pH} + \text{pOH} = 14.000$
 $\text{pH} + 5.640 = 14.000$
 $\therefore \text{pH} = 14.000 - 5.640 = 8.360$ \rightarrow 3 decimal places
 $\therefore [\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-8.360}$
3 sig figs $\leftarrow 4.365158322 \times 10^{-9} \text{ M}$
4.37 $\times 10^{-9} \text{ M}$

35. The K_b of ethylamine is 4.30×10^{-4} . What is the pH of a 0.0847 M aqueous solution of ethylamine?
- constant is not included in sig fig calculation*

11.781

D)

	$\text{CH}_3\text{CH}_2\text{NH}_2$	$[\text{H}_2\text{O}]$	$\text{CH}_3\text{CH}_2\text{NH}_3^+$	$[\text{OH}^-]$
Initial	0.0847	—	0	0
Change	$-x$	—	$+x$	$+x$
Equilibrium	$0.0847-x$	—	x	x

3 sig figs
∴ pH should have 3 decimal places

$$K_b = \frac{[\text{CH}_3\text{CH}_2\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{CH}_2\text{NH}_2]}$$

$$\therefore 4.30 \times 10^{-4} = \frac{[x][x]}{[0.0847-x]}$$

$\frac{C}{K} \gg 100$

$$\therefore x^2 = 4.30 \times 10^{-4} (0.0847)$$

$$\therefore x = \pm 6.034981359 \times 10^{-3}$$

ignore the negative value

$$[\text{OH}^-] = x = 6.03 \times 10^{-3} \text{ M}$$

$$\therefore \text{pOH} = -\log [\text{OH}^-] = -\log [6.03 \times 10^{-3}] = 2.219$$

$$\text{pH} = 14.000 - 2.219 = 11.781$$

* Note: do not round until the end

36. Vinegar is a 5.0% solution by weight of acetic acid ($\text{CH}_3\text{CO}_2\text{H}$) (60.05 g/mol) in water. Given that $K_a = 1.8 \times 10^{-5}$ for acetic acid and assuming the density of vinegar to be 1.00 g/cm³, what is the pH of this vinegar solution?
- 2 sig figs*

2.41

	$[\text{CH}_3\text{CO}_2\text{H}]$	$[\text{H}_2\text{O}]$	$[\text{CH}_3\text{CO}_2^-]$	$[\text{H}_3\text{O}^+]$
Initial	0.83	—	0	0
Change	$-x$	—	$+x$	$+x$
Equilibrium	$0.83-x$	—	x	x

Determine the initial conc. of $\text{CH}_3\text{CO}_2\text{H}$ in $\frac{\text{mol}}{\text{L}}$:

$$\frac{5.0 \text{ g CH}_3\text{COOH}}{100.0 \text{ g solution}} \times \frac{1 \text{ mol CH}_3\text{COOH}}{60.05 \text{ g CH}_3\text{COOH}} \times \frac{1.00 \text{ g}}{1 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.8326894671 \text{ M}$$

$$K_a = \frac{[\text{CH}_3\text{CO}_2^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{CO}_2\text{H}]}$$

$$\therefore 1.8 \times 10^{-5} = \frac{[x][x]}{[0.83-x]}$$

$$\therefore x^2 = 1.85 \times 10^{-5} [0.83]$$

$$\therefore x^2 = 1.5 \times 10^{-5}$$

$$\therefore x = \pm 3.871370611 \times 10^{-3}$$

$$[\text{H}_3\text{O}^+] = 3.9 \times 10^{-3} \text{ M}$$

$$\therefore \text{pH} = -\log [\text{H}_3\text{O}^+] = -\log [3.9 \times 10^{-3}] = 2.41 \sim 2 \text{ decimal places}$$

37. At 50 °C the value of K_w is 5.50×10^{-14} . A basic solution at 50 °C has

B

- A. $[\text{H}_3\text{O}^+] < [\text{OH}^-] < 2.35 \times 10^{-7} \text{ M}$.
 B. $[\text{H}_3\text{O}^+] < 2.35 \times 10^{-7} \text{ M} < [\text{OH}^-]$.
 C. $[\text{H}_3\text{O}^+] = [\text{OH}^-] < 2.35 \times 10^{-7} \text{ M}$.
 D. $[\text{H}_3\text{O}^+] > [\text{OH}^-] > 2.35 \times 10^{-7} \text{ M}$.

K_w @ 25°C
 1.0×10^{-14}

1) $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 5.50 \times 10^{-14}$

2) If a solution is basic

$[\text{H}_3\text{O}^+] = [\text{OH}^-] = x$

$\therefore [\text{OH}^-] > 2.35 \times 10^{-7} \text{ M}$

$\therefore [x][x] = 5.50 \times 10^{-14}$

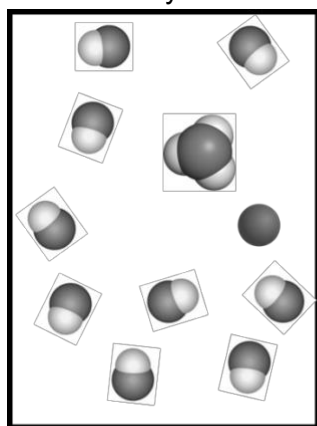
$x^2 = 5.50 \times 10^{-14}$

$\therefore x = \pm 2.34520788 \times 10^{-7}$

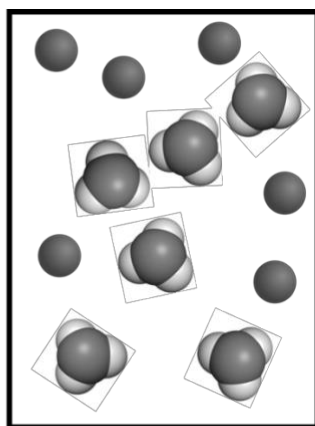
K

$[\text{H}_3\text{O}^+] < 2.35 \times 10^{-7} \text{ M}$

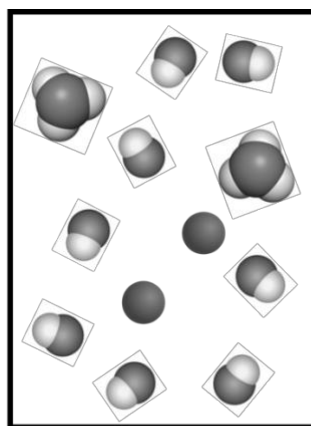
38. The pictures represent aqueous solutions of three acids HA (A = X, Y, or Z); water molecules have been omitted for clarity.



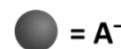
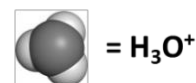
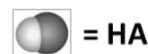
HX



HY



HZ



Which acid has the **smallest** value of K_a ?

A

- A. HX
 B. HY
 C. HZ

D. All have the same K_a value.

- K_a is a measure of the strength of weak acids

- K_a is also a measure of the extent of ionization

- the lower the K_a the less ionized the acid will be & vice versa
 heat +

$K_w = 1.0 \times 10^{-14}$ at 25°C

* If water is heated by raising its temperature above 25°C, the reaction will shift to the right increasing the equilibrium amounts of H_3O^+ & OH^- .
 this is due to water being an exothermic process.

* If the amounts of H_3O^+ & OH^- the value of K_w increases
 $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$

→ HA

39. A monoprotic acid, Para-aminobenzoic acid (PABA), $p\text{-H}_2\text{NC}_6\text{H}_4(\text{COOH})$, is used in some sunscreens and hair conditioning products. Calculate the pH of an aqueous solution with $[\text{PABA}] = 0.030 \text{ M}$ and $K_a = 2.2 \times 10^{-5}$.

- A. 1.52
B. 3.09
C. 4.66
D. 6.18

B



	$[\text{HA}]$	$[\text{H}_2\text{O}]$	$[\text{A}^-]$	$[\text{H}_3\text{O}^+]$
Initial	0.030	—	0	0
Change	$-x$	—	$+x$	$+x$
Equilibrium	$0.030 - x$	—	x	x

2) $K_a = \frac{[\text{A}^-][\text{H}_3\text{O}^+]}{[\text{HA}]}$

$2.2 \times 10^{-5} = \frac{[x][x]}{[0.030 - x]}$

$\frac{C}{K} > 100$

$\therefore x^2 = 2.2 \times 10^{-5} [0.030] = 6.6 \times 10^{-7}$

$\therefore x = \pm 8.124038405 \times 10^{-4}$ → disregard the negative value

$\therefore \text{pH} = -\log [8.1 \times 10^{-4}] = 3.09$ → 2 decimal places

2 sig figs

40. Given the acids and their K_a values:

Hydrocyanic acid, HCN

$K_a = 4.00 \times 10^{-10}$

Phenol, $\text{C}_6\text{H}_5\text{OH}$

$K_a = 1.00 \times 10^{-10}$

Benzoic acid, $\text{C}_6\text{H}_5\text{CO}_2\text{H}$

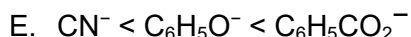
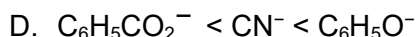
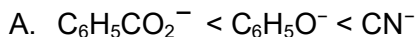
$K_a = 6.30 \times 10^{-5}$

weakest to strongest

K_a is a reflection of acid strength

What is the order of increasing base strength for CN^- , $\text{C}_6\text{H}_5\text{O}^-$, and $\text{C}_6\text{H}_5\text{CO}_2^-$?

D



* the higher the K_a of an acid, the stronger the acid. A stronger acid produces a weak conjugate base & the weaker the acid the stronger the conjugate base

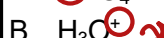
* order of acid strength in increasing order (weakest to strongest):



* order of conjugate base strength in increasing order: $\text{C}_6\text{H}_5\text{CO}_2^- < \text{CN}^- < \text{C}_6\text{H}_5\text{O}^-$

41. Which of these species is amphoteric?

A



E. None of the above are amphoteric

has a negative charge... it can accept an electron

has a proton that it can donate

has an extra proton \therefore an acid

has a negative charge only \therefore base

same as C

An amphoteric species is capable of both donating a proton and accepting a proton

heat +

$$\approx K_w = 1.0 \times 10^{-14} \text{ @ } 25^\circ\text{C}$$

42. Given: $\text{H}_2\text{O(l)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$, $\Delta H^\circ_{\text{rxn}} > 0$ When the temperature of a sample of pure water is raised above 25°C ,

C

- A. the hydronium ion concentration will be greater than the hydroxide ion concentration.
- B. the hydronium ion concentration will be less than the hydroxide ion concentration.
- C. the value of K_w will increase.
- D. the hydronium ion concentration could change to $1.0 \times 10^{-10} \text{ M}$.
- E. the hydroxide ion concentration could change to $1.0 \times 10^{-10} \text{ M}$.

* If water is heated by raising its temperature above 25°C , the reaction will shift to the right increasing the equilibrium amounts of H_3O^+ & OH^- .
this is due to water being an exothermic process.

* If the amounts of H_3O^+ & OH^- the value of K_w increases

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

43. The magnitude of K_w indicates that _____.

C

- A. water autoionizes very slowly
- B. water autoionizes very quickly
- C. water autoionizes only to a very small extent
- D. the autoionization of water is exothermic

} Equilibrium constants indicate extents

K_w is the equilibrium constant for the auto ionization of water

44. Which of the following acids has the lowest pH?

A

- A. 0.1 M HA, $pK_a = 2.43$
- B. 0.1 M HB, $pK_a = 4.55$
- C. 0.1 M HC, $pK_a = 8.23$
- D. 0.1 M HD, $pK_a = 11.89$

* All the acids have the same concentration
 \therefore the pH will be directly related to pK_a
the stronger the acid the $\uparrow K_a$ & the $\downarrow pK_a$
 $\therefore \downarrow \text{pH}$

45. For which of the following solutions must the ionization of water be considered when calculating the pH or pOH? Select all that apply.

A D

A. $3 \times 10^{-8} \text{ M HNO}_3$

~~B.~~ 0.10 g HCl in 1.0 L of solution

$2.8 \times 10^{-3} \text{ M HCl}$

~~C.~~ 0.00080 g NaOH in 0.50 L of solution

$4.0 \times 10^{-5} \text{ M NaOH}$

D. $1 \times 10^{-7} \text{ M Ca(OH)}_2$

the autoionization of water must be considered in the pH calculations of an extremely dilute strong acid or strong base solution ($< 10^{-6} \text{ M}$)

46. Phenylacetic acid ($\text{C}_6\text{H}_5\text{CH}_2\text{COOH}$, simplified here as HPAC) builds up in the blood of persons with phenylketonuria, an inherited disorder that, if untreated, causes mental retardation and death. A study of the acid shows that the pH of 0.12 M HPAC is 2.62. What is the pK_a of phenylacetic acid?

4.31

* Also 4.32 is correct



	$[\text{HPAC}]$	$[\text{H}_2\text{O}]$	$[\text{H}_3\text{O}^+]$	$[\text{PAC}^-]$
Initial	0.12	—	0	0
Change	-2.4×10^{-3}	—	2.4×10^{-3}	2.4×10^{-3}
Equilibrium	0.12	—	2.4×10^{-3}	2.4×10^{-3}

found from pH

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-2.62} = 2.4 \times 10^{-3}$$

$$2) K_a = \frac{[\text{H}_3\text{O}^+][\text{PAC}^-]}{[\text{HPAC}]} = \frac{[2.4 \times 10^{-3}][2.4 \times 10^{-3}]}{[0.12]}$$

$$= 4.893148185 \times 10^{-5}$$

→ 2 sig fig

$$3) \text{ Calculate } pK_a = -\log K_a = -\log (4.893148185 \times 10^{-5})$$

$$= 4.310411632$$

$$\approx 4.31 \rightarrow 2 \text{ decimal places}$$

Formula Sheet

Length

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

Energy

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

Pressure

1 pascal = $1 \text{ N} / \text{m}^2 = 1 \text{ kg} / \text{m} \cdot \text{s}^2$
1 atmosphere = 101.325 kilopascals = 760 mm Hg = 760 torr = 14.70 lb/in²
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 = $1 \text{ cm}^3 = 1 \text{ cc}$

Constants

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

Equations

$d \text{ (density)} = m/V$
 $P_1 V_1 = P_2 V_2$
 $V_1/T_1 = V_2/T_2$
 $P_1 V_1/n_1 T_1 = P_2 V_2/n_2 T_2$
 $PV = nRT$
 $(P + a(n^2/V^2)) \cdot (V - nb) = nRT$
molar mass (M) = mRT/PV
density (d) = MP/RT
 $x_A = n_A/n_{\text{tot}} = P_A/P_{\text{tot}} = V_A/V_{\text{tot}}$
 $P_{\text{tot}} = P_A + P_B + \dots$
 $n_{\text{tot}} = n_A + n_B + \dots$
$$\mu_{rms} = \sqrt{\frac{3RT}{M}}$$

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

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

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

$$w = -P\Delta V$$

$$\Delta E = q + w$$

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

$$\Delta H^\circ = \sum n\Delta H^\circ(\text{bonds broken}) - \sum n\Delta H^\circ(\text{bonds formed})$$

$$E = h\nu$$

$$c = \lambda\nu$$

$$\lambda = h/mv$$

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

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

$$C_g = kP_g$$

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

$$P_{\text{solution}} = \sum P_j = \sum P_j X_j$$

$$\Delta T_b = K_b m_i$$

$$\Delta T_f = K_f m_i$$

$$\pi = MRTi$$

Thermodynamic and Electrochemistry

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

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

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

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

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

$$\Delta S^\circ_{\text{rxn}} = \sum \nu S^\circ_{\text{products}} - \sum \nu S^\circ_{\text{reactants}}$$

$$\Delta H^\circ_{\text{rxn}} = \sum \nu H^\circ_{\text{products}} - \sum \nu H^\circ_{\text{reactants}}$$

$$\Delta G^\circ_{\text{rxn}} = \sum \nu G^\circ_{\text{products}} - \sum \nu G^\circ_{\text{reactants}}$$

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

$$\Delta G = \Delta G^\circ + RT \cdot \ln Q$$

$$R = 8.314 \text{ J/mol.K}$$

$$\Delta G^\circ = -RT \cdot \ln K$$

$$\Delta G = -nFE_{\text{cell}}$$

$$F = 96485 \text{ J/(V}\cdot\text{mol e}^-)$$

$$E^\circ_{\text{cell}} = RT/nF \ln K$$

$$E^\circ_{\text{cell}} = (0.0257/n) \ln K = (0.0592/n) \log K$$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - (RT/nF) \ln Q$$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - (0.0257/n) \ln Q$$

$$\text{Electrolysis: } Q \text{ (total charge)} = I \times t = n \times F$$

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

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

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$K_w = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

$$K_w = [\text{H}_3\text{O}^+] \times [\text{OH}^-]$$

$$K_w = K_a \times K_b$$

$$\text{p}K_a = -\log[K_a]$$

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

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

1

1 H 1.01

3 Li 6.94	4 Be 9.01
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11 Na 22.99	12 Mg 24.31
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19 K 39.10	20 Ca 40.08	21 Sc 44.96
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37 Rb 85.47	38 Sr 87.62	39 Y 88.91
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37 Cs 132.91	56 Ba 137.33
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87 Fr [223]	88 Ra [226]
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18

2 He 4.00

5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
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13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
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31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80
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49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
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81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po [209]	85 At [210]	86 Rn [222]
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113 Nh [286]	114 Fl [290]	115 Mc [290]	116 Lv [293]	117 Ts [294]	118 Og [294]
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Periodic Table of the Elements

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]