

## Recitation Worksheet Ten: Exam 3 Review

Name:

*Key*

MyID:

### Textbook:

Chemistry & Chemical Reactivity

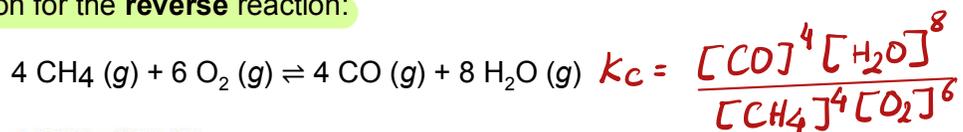
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### Instructions:

- 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 28<sup>th</sup>– 31<sup>st</sup>) 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

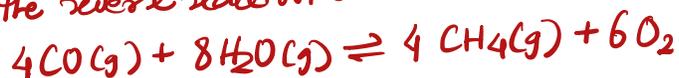
A.  $K_c' = \frac{4[\text{CO}] + 8[\text{H}_2\text{O}]}{4[\text{CH}_4] + 6[\text{O}_2]}$

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:

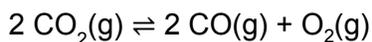


$K_c$  for this reaction

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

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



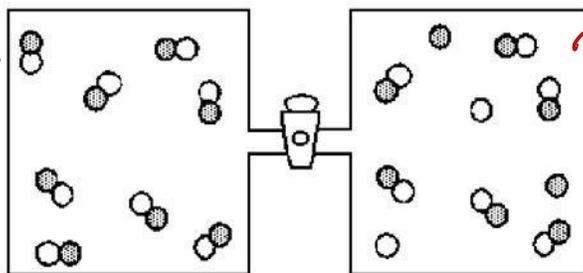
$$K_p = \frac{P(\text{CO})^2 P(\text{O}_2)}{P(\text{CO}_2)^2}$$

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

$$= 0.01890432099$$
$$\approx 0.019$$

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  $\therefore K_c = \frac{[AB]^2}{[A][B]}$   
 $= \frac{[6]^2}{[2][2]} = \frac{36}{4} = 9$

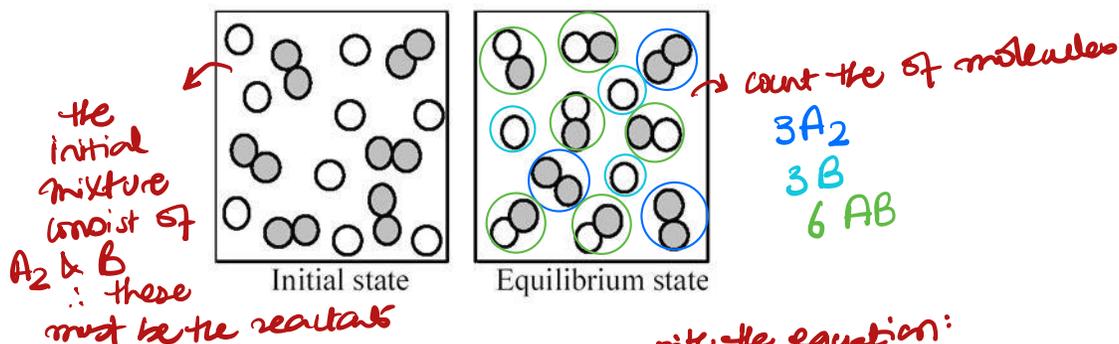
If the barrier between the two vessels is removed and the contents of the two vessels are allowed to mix, what will be observed?

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.

$Q_c = \frac{[AB]^2}{[A][B]} = \frac{[8+6]^2}{[2][2]} = 49$   $\rightarrow$  From the vessel to the left  $\therefore Q_c > K_c$  & 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.  $A_2 + 2B \rightleftharpoons A_2B_2$   
 C.  $A_2 + 2B \rightleftharpoons 2AB$   
 D.  $6A_2 + 9B \rightleftharpoons 3A_2 + 3B + 6AB$

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

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.

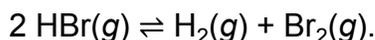
↳ Altering the temperature of the reaction changes the value of  $K_c$

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  
 $\therefore K_c = \frac{[\text{B}][\text{C}]}{[\text{A}]^2}$  ↑

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  $\text{H}_2$  at equilibrium?

A

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

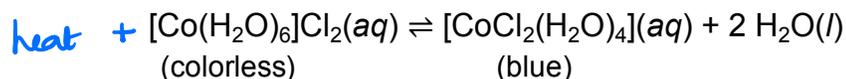
	[HBr]	[H <sub>2</sub> ]	[Br <sub>2</sub> ]
Initial	0.300	0	0
Change	-2x	+0.0520	+x
Equilibrium	0.196	0.0520	0.0520

$$K_c = \frac{[\text{H}_2][\text{Br}_2]}{[\text{HBr}]^2}$$

$$= \frac{[0.0520][0.0520]}{[0.196]^2}$$

$$= \frac{0.002704}{0.038416} = 0.0703873386 \approx 7.04 \times 10^{-2}$$

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

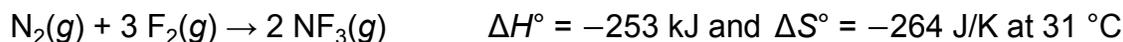


If the reactant solution is used to write on a piece of paper and the paper is allowed to partially dry, what can be done to bring out the colored handwriting?

- 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  $\Delta G^\circ$  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.

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

$$\begin{aligned} \text{Also } \Delta G^\circ &= -RT \ln K \\ \text{if } \Delta G^\circ < 0 &\therefore K > 1 \\ \Delta G^\circ > 0 &K < 1 \\ \Delta G^\circ = 0 &K = 1 \end{aligned}$$

*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.

Increasing time →

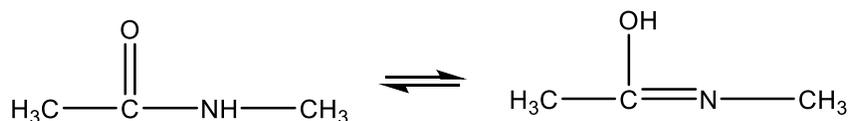
(1)					$k_c = \frac{9}{3} = 3$  $k_c = \frac{7}{5} = 1.4$  $k_c = 3$  $k_c = 3$
(2)					
(3)					
(4)					

From the last (left to right) picture you can calculate  $k_c = \frac{[B]}{[A]}$

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

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

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

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

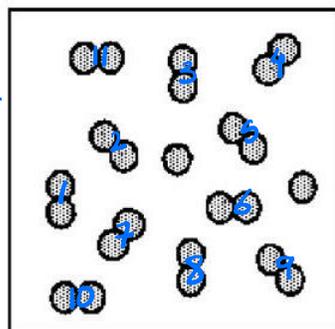
$\therefore Q_C > K_C$  & the reaction shifts to the left (towards reactants)

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

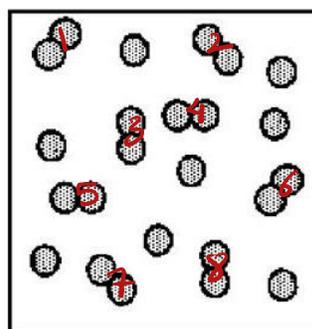
13. Consider the reaction  $2A(g) \rightleftharpoons A_2(g)$ . The pictures represent the initial state and the equilibrium state of the system.

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  $\Delta G$  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

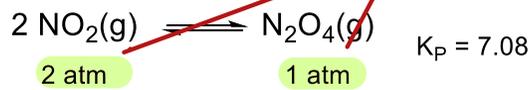
C

- A.  $\Delta G$  and  $\Delta G^\circ$  increase.  
 B.  $\Delta G$  and  $\Delta G^\circ$  decrease.  
 C.  $\Delta G$  increases, but  $\Delta G^\circ$  remains constant.  
 D.  $\Delta G$  decreases, but  $\Delta G^\circ$  remains constant.

*At a constant temperature there is only one  $\Delta G^\circ$  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 &  $\Delta G$  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*

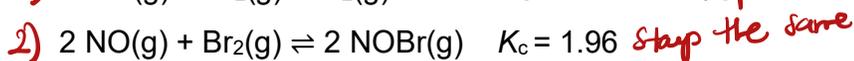
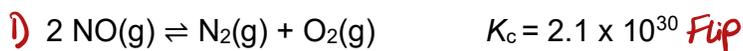
*C) if  $K_p > 1 \therefore k_f > k_r \therefore$  incorrect*

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

*$K = \frac{k_{f \rightarrow \text{rate}} \text{ and not forward} > 1}{k_{r \rightarrow \text{rate}} \text{ and not for the reverse process}}$*

*$K_p = 7.08 > 1$*

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



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

0

A.  $9.7 \times 10^{-16}$

B.  $2.1 \times 10^{30}$

C.  $4.1 \times 10^{30}$

D.  $9.3 \times 10^{-31}$

E.  $8.7 \times 10^{-61}$



$\therefore K_c = 9.33333333 \times 10^{-31}$   
 $\approx 9.3 \times 10^{-31}$

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

$3.82 \times 10^{-32}$

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

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

$= 3.81670857 \times 10^{-32}$

$\approx 3.82 \times 10^{-32}$

$\therefore$  infinit sig figs

constant

2 sig figs based on temperature

19. What is the value of  $\Delta G^\circ$  in kJ/mol for the reaction  $2 \text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2 \text{SO}_3(g)$  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

$\Delta G^\circ = -RT \ln K_p$   $\rightarrow$  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  $\Delta G^\circ$

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

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  $K_c$

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

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

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

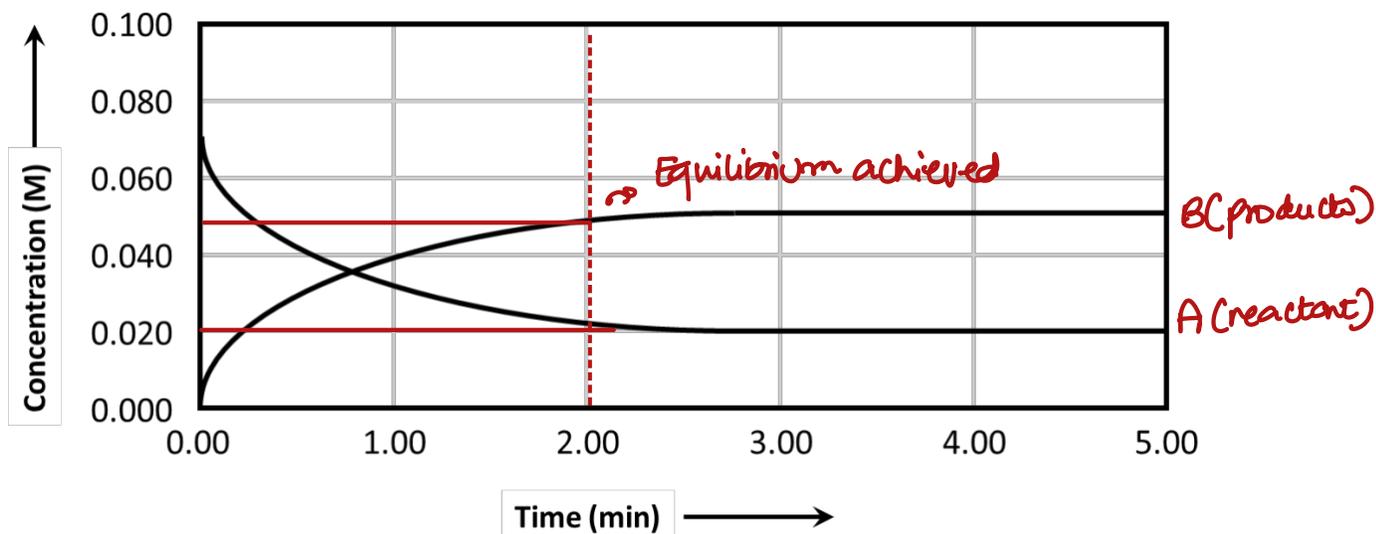
$$0.20 = 3x$$

$$\therefore x = 0.067$$

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



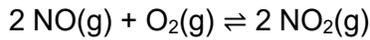
**D**

- 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 O<sub>2</sub> are present at equilibrium in a 2.50 L flask. If the value for K<sub>c</sub> is 23.8, how many moles of NO<sub>2</sub> 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: N<sub>2</sub>O<sub>4</sub>(g) ⇌ 2 NO<sub>2</sub>(g) for which K<sub>p</sub> = 0.1134 at 25 °C and ΔH<sub>r</sub><sup>°</sup> = 58.03 kJ/mol. Assume that 1 mole of N<sub>2</sub>O<sub>4</sub> and 2 moles of NO<sub>2</sub> are introduced into a 5.0-liter container. Calculate the equilibrium value of [N<sub>2</sub>O<sub>4</sub>]?

0.379

M

	[N <sub>2</sub> O <sub>4</sub> ]	[NO <sub>2</sub> ]
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<sub>c</sub> from K<sub>p</sub>

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

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

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

Q<sub>c</sub> > K<sub>c</sub> ∴ reaction proceeds in the reverse direction

3) Calculate x by plugging K<sub>c</sub>

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

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

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

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

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

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

disregard this value

24. Consider the equilibria:

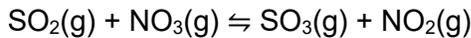


$$K_c = 2.3 \times 10^{-7}$$



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

Calculate the equilibrium constant for the reaction



- A**
- A. 78
  - B.  $1.3 \times 10^{-2}$
  - C.  $1.6 \times 10^{-4}$
  - D.  $3.2 \times 10^{-10}$
  - E.  $6.1 \times 10^3$

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  $9.3 \times 10^9$  at 298 K and  $3.3 \times 10^7$  at 398 K. Calculate:

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

**-56.9** kJ/mol

$\Delta G^\circ_{\text{rxn}} = -RT \ln K_p$   $\rightarrow$  sig fig 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.  $\Delta H^\circ$  between 298 K and 398 K

**-55.6** kJ/mol

$\rightarrow$  sig fig are based on temperature

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

$K_1 = 9.3 \times 10^9$   $R = 8.314 \times 10^{-3} \text{ kJ/mol}\cdot\text{K}$

$T_1 = 298$   $\Delta H^\circ = ?$

$K_2 = 3.3 \times 10^7$

$T_2 = 398$

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

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

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

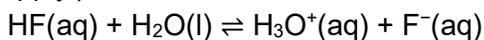
26. Which of these pairs is **NOT** a Brønsted-Lowry conjugate acid-base pair?

**B**

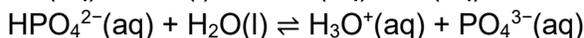
- A. HOBr and OBr<sup>-</sup>
- B. H<sub>2</sub>SO<sub>4</sub> and SO<sub>4</sub><sup>2-</sup> → difference is two protons
- C. C<sub>6</sub>H<sub>5</sub>NH<sub>3</sub><sup>+</sup> and C<sub>6</sub>H<sub>5</sub>NH<sub>2</sub>
- D. H<sub>2</sub>CO<sub>3</sub> and HCO<sub>3</sub><sup>-</sup>
- E. C<sub>6</sub>H<sub>5</sub>CH<sub>2</sub>CO<sub>2</sub>H and C<sub>6</sub>H<sub>5</sub>CH<sub>2</sub>CO<sub>2</sub><sup>-</sup>

The difference between a Brønsted-Lowry conjugate acid-base pair is only one proton (H<sup>+</sup>)

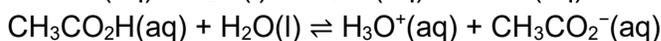
27. For these three acids and their acid equilibrium constants, which of the statements are true? (Select all that apply.)



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



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



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

**BDE**

- A. HPO<sub>4</sub><sup>2-</sup> is the strongest acid *False: has the smallest equilibrium constant*
- B. HF produces the weakest conjugate base out of the three acids *True: HF is the strongest acid out of the three acids ∴ produces the weakest conjugate base*
- C. The rate of dissociation of HF is the fastest as indicated by its acid equilibrium constant *False: K<sub>a</sub> gives information about the extent of the dissociation of an acid*
- D. The conjugate base of CH<sub>3</sub>CO<sub>2</sub>H is CH<sub>3</sub>CO<sub>2</sub><sup>-</sup> *True*
- E. HF is the strongest acid *True (largest K<sub>a</sub>)*

28. You are given four solutions A, B, C, and D. Solution A has an [H<sub>3</sub>O<sup>+</sup>] = 1.7 × 10<sup>-4</sup> M, solution B has a pH = 5.25, solution C has an [OH<sup>-</sup>] = 1.5 × 10<sup>-5</sup> and solution D has a pOH = 1.55. Arrange the solutions in the order of increasing [H<sub>3</sub>O<sup>+</sup>].

smallest to largest

**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: [H<sub>3</sub>O<sup>+</sup>] = 1.7 × 10<sup>-4</sup> M

B) Solution B: pH = 5.25

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-5.25} = 5.6 \times 10^{-6} \text{ M}$$

C) Solution C: [OH<sup>-</sup>] = 1.5 × 10<sup>-5</sup>

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$\therefore [\text{H}_3\text{O}^+][1.5 \times 10^{-5}] = 1.0 \times 10^{-14}$$

$$\therefore [\text{H}_3\text{O}^+] = 6.7 \times 10^{-10}$$

D) Solution D: pOH = 1.55

$$\therefore \text{pH} + \text{pOH} = 14.00$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 1.55$$

$$= 12.45$$

$$[\text{H}_3\text{O}^+] = 10^{-12.45} = 3.5 \times 10^{-13} \text{ M}$$

Solutions in increasing order of [H<sub>3</sub>O<sup>+</sup>]:

D < C < B < A

D < C < B < A

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.  $2.9 \times 10^{-4}$
- B.  $4.4 \times 10^{-6}$
- C.  $8.7 \times 10^{-8}$
- D.  $1.7 \times 10^{-11}$
- E.  $1.1 \times 10^{-6}$

i)



	[HA]	[H <sub>2</sub> O]	[A <sup>-</sup> ]	[H <sub>3</sub> O <sup>+</sup> ]
Initial	$2.0 \times 10^{-2}$	—	0	0
Change	$-3.0 \times 10^{-4}$	—	$+3.0 \times 10^{-4}$	$+3.0 \times 10^{-4}$
Equilibrium	$2.0 \times 10^{-2}$	—	$3.0 \times 10^{-4}$	$3.0 \times 10^{-4}$

→ [H<sub>3</sub>O<sup>+</sup>]<sub>eq</sub> can be determined from pH  
 $\therefore [H_3O^+]_{eq} = 10^{-pH} = 10^{-3.53} = 3.0 \times 10^{-4}$   
 2 decimal places

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

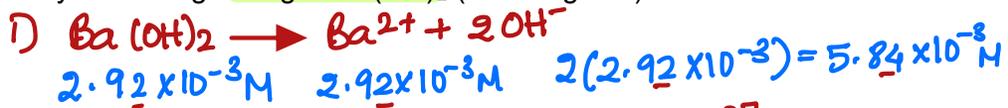
\* Note: Do not round until the end

30. What is the pH of a solution prepared by dissolving 1.50 g of Ba(OH)<sub>2</sub> (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

→ strong base



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

$$3) pH + pOH = 14.000$$

$$\therefore pH = 14.000 - pOH = 14.000 - 2.234 = 11.766 \rightarrow 3 \text{ decimal places}$$

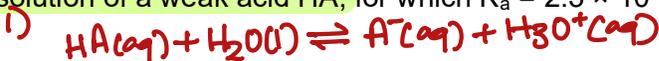
Molarity of Ba(OH)<sub>2</sub>:

$$\frac{1.50 \text{ g Ba(OH)}_2 \times \frac{1 \text{ mol Ba(OH)}_2}{171.34 \text{ g Ba(OH)}_2}}{3.00 \text{ L}} = 2.91817429 \times 10^{-3} M \text{ (3 sig fig)}$$

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 <sub>2</sub> O]	[A <sup>-</sup> ]	[H <sub>3</sub> O <sup>+</sup> ]
Initial	0.135	-	0	0
Change	-x	-	+x	+x
Equilibrium	0.135-x	-	x	x

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

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

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

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

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

$\therefore x = 0.04539688647$

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

or  $x = -0.068896886 \rightarrow$

disregard this value

3)  $[H_3O^+] =$

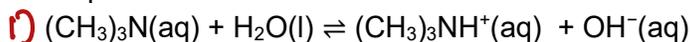
$\therefore pH = -\log [H_3O^+]$

$= -\log [0.0454]$

$= 1.343$

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 <sub>3</sub> ) <sub>3</sub> N]	[H <sub>2</sub> O]	[(CH <sub>3</sub> ) <sub>3</sub> NH <sup>+</sup> ]	[OH <sup>-</sup> ]
Initial	x	-	0	0
Change	-3.2 x 10 <sup>-4</sup>	-	+3.2 x 10 <sup>-4</sup>	+3.2 x 10 <sup>-4</sup>
Equilibrium	x - 3.2 x 10 <sup>-4</sup>	-	3.2 x 10 <sup>-4</sup>	3.2 x 10 <sup>-4</sup>

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

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

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

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

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

$\therefore [(CH_3)_3N]_{eq} = 1.903529353 \times 10^{-3} - 3.16227766 \times 10^{-4} = 1.587301687 \times 10^{-3} \approx 1.6 \times 10^{-3} M$

33. What is the conjugate base of sulfuric acid?

A

- A.  $\text{HSO}_4^-$
- B.  $\text{H}_3\text{O}^+$
- C.  $\text{OH}^-$
- D.  $\text{SO}_4^{2-}$
- E.  $\text{H}_3\text{SO}_4^+$

$\text{H}_2\text{SO}_4$  (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  $25^\circ\text{C}$  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.  $8.360 \text{ M}$

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

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

11.781

$$\text{CH}_3\text{CH}_2\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{CH}_2\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$$

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

3 sig figs  
 $\therefore$  pH should have 3 decimal places

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

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

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

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

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

$\hookrightarrow$  disregard 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/cm<sup>3</sup>, 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  $\frac{\text{mol}}{\text{L}}$ :

$$\frac{5.0 \text{ g CH}_3\text{CO}_2\text{H}}{100.0 \text{ g solution}} \times \frac{1 \text{ mol CH}_3\text{CO}_2\text{H}}{60.05 \text{ g CH}_3\text{CO}_2\text{H}} \times \frac{1.00 \text{ g}}{1 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.8326894671 \text{ M}$$

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

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

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

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

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

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

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

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

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

**B**

- A.  $[H_3O^+] < [OH^-] < 2.35 \times 10^{-7} M$ .
- B.  $[H_3O^+] < 2.35 \times 10^{-7} M < [OH^-]$ .
- C.  $[H_3O^+] = [OH^-] < 2.35 \times 10^{-7} M$ .
- D.  $[H_3O^+] > [OH^-] > 2.35 \times 10^{-7} M$ .

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

1)  $K_w = [H_3O^+][OH^-] = 5.50 \times 10^{-14}$

$[H_3O^+] = [OH^-] = x$

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

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

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

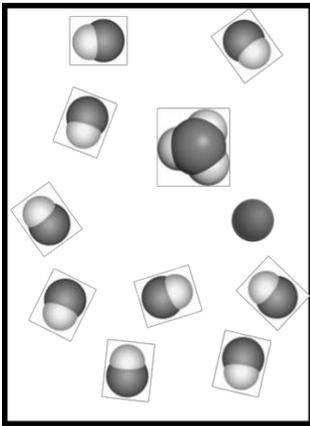
2) If a solution is basic

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

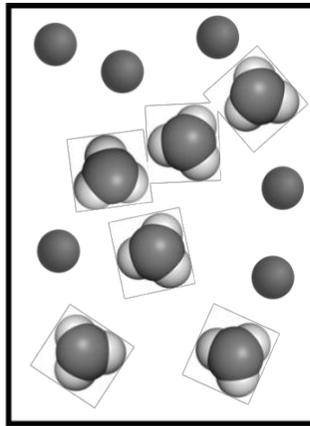
$K$

$[H_3O^+] < 2.35 \times 10^{-7} M$

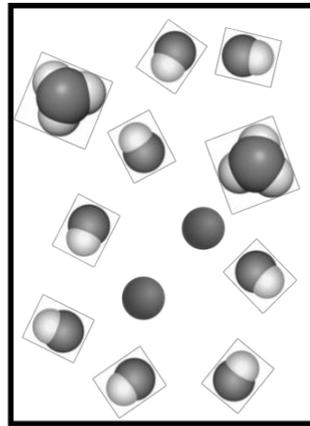
38. The pictures represent aqueous solutions of three acids HA (A = X, Y, or Z); water molecules have been omitted for clarity.



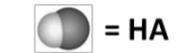
HX



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  $25^\circ C$

\* If water is heated by raising its temperature above  $25^\circ C$ , 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  
 $K_w = [H_3O^+][OH^-]$

→ 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	$0.030 - x$	-	$x$	$x$

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

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

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

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

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

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

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  $\text{CN}^-$ ,  $\text{C}_6\text{H}_5\text{O}^-$ , and  $\text{C}_6\text{H}_5\text{CO}_2^-$ ?

D

- A.  $\text{C}_6\text{H}_5\text{CO}_2^- < \text{C}_6\text{H}_5\text{O}^- < \text{CN}^-$
- B.  $\text{C}_6\text{H}_5\text{O}^- < \text{C}_6\text{H}_5\text{CO}_2^- < \text{CN}^-$
- C.  $\text{CN}^- < \text{C}_6\text{H}_5\text{CO}_2^- < \text{C}_6\text{H}_5\text{O}^-$
- D.  $\text{C}_6\text{H}_5\text{CO}_2^- < \text{CN}^- < \text{C}_6\text{H}_5\text{O}^-$
- E.  $\text{CN}^- < \text{C}_6\text{H}_5\text{O}^- < \text{C}_6\text{H}_5\text{CO}_2^-$

\* 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

- A.  $\text{HPO}_4^{2-}$  has a negative charge... it can accept an electron
- B.  $\text{H}_3\text{O}^+$  has an extra proton  $\therefore$  an acid
- C.  $\text{PO}_4^{3-}$  has a negative charge only  $\therefore$  base
- D.  $\text{Cl}^-$  same as C
- E. None of the above are amphoteric

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

heat +

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

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

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  $25^\circ\text{C}$ , the reaction will shift to the right increasing the equilibrium amounts of  $\text{H}_3\text{O}^+$  &  $\text{OH}^-$ . This is due to water being an exothermic process.

\* If the amounts of  $\text{H}_3\text{O}^+$  &  $\text{OH}^-$  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 extent

$K_w$  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.  $3 \times 10^{-8}$  M HNO<sub>3</sub>
- ~~B.~~ 0.10 g HCl in 1.0 L of solution  $2.8 \times 10^{-3}$  M HCl
- ~~C.~~ 0.00080 g NaOH in 0.50 L of solution  $4.0 \times 10^{-5}$  M NaOH
- D.  $1 \times 10^{-7}$  M Ca(OH)<sub>2</sub>

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

46. Phenylacetic acid (C<sub>6</sub>H<sub>5</sub>CH<sub>2</sub>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<sub>a</sub> of phenylacetic acid?

4.31

\* Also 4.32 is correct



	[HPAc]	[H <sub>2</sub> O]	[H <sub>3</sub> O <sup>+</sup> ]	[PAC <sup>-</sup> ]
Initial	0.12	-	0	0
Change	$-2.4 \times 10^{-3}$	-	$2.4 \times 10^{-3}$	$2.4 \times 10^{-3}$
Equilibrium	0.12	-	$2.4 \times 10^{-3}$	$2.4 \times 10^{-3}$

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

2)  $K_a = \frac{[\text{H}_3\text{O}^+][\text{PAC}^-]}{[\text{HPAc}]} = \frac{[2.4 \times 10^{-3}][2.4 \times 10^{-3}]}{[0.12]}$   
 $= 4.893148185 \times 10^{-5}$

3) Calculate  $\text{p}K_a = -\log K_a = -\log(4.893148185 \times 10^{-5})$   
 $= 4.310411632$   
 $\approx 4.31$  (2 decimal places)

## Formula Sheet

### Length

1 kilometer = 0.62137 mile  
1 inch = 2.54 centimeters (exactly)  
1 Ångstrom =  $1 \times 10^{-10}$  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/in<sup>2</sup>  
1 bar =  $1 \times 10^5$  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

$c = 2.998 \times 10^8 \text{ m}/\text{sec}$   
 $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{sec}^{-1}$   
 $R = 0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K} = 8.314 \text{ J}/\text{mol}\cdot\text{K}$   
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:  $6.022 \times 10^{23}$

### Equations

d (density) =  $m/V$

$P_1V_1 = P_2V_2$

$V_1/T_1 = V_2/T_2$

$P_1V_1/n_1T_1 = P_2V_2/n_2T_2$

$PV = nRT$

$(P + a(n^2/V^2))\cdot(V - nb) = nRT$

molar mass (M) =  $mRT/PV$

density (d) =  $MP/RT$

$x_A = n_A/n_{\text{tot}} = P_A/P_{\text{tot}} = V_A/V_{\text{tot}}$

$P_{\text{tot}} = P_A + P_B + \dots$

$n_{\text{tot}} = n_A + n_B + \dots$

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

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

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

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

$$w = -P\Delta V$$

$$\Delta E = q + w$$

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

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

$$E = hv$$

$$c = \lambda\nu$$

$$\lambda = h/mv$$

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

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

$$C_g = kP_g$$

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

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

$$\Delta T_b = K_b m_i$$

$$\Delta T_f = K_f m_i$$

$$\pi = MRT_i$$

### Thermodynamic and Electrochemistry

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

### 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

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

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

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

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

$$K_w = K_a \times K_b$$

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

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

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

# Periodic Table of the Elements

1 <b>H</b> 1.01																	2 <b>He</b> 4.00									
3 <b>Li</b> 6.94	4 <b>Be</b> 9.01											5 <b>B</b> 10.81	6 <b>C</b> 12.01	7 <b>N</b> 14.01	8 <b>O</b> 16.00	9 <b>F</b> 19.00	10 <b>Ne</b> 20.18									
11 <b>Na</b> 22.99	12 <b>Mg</b> 24.31	3 <b>Al</b> 26.98	13 <b>Si</b> 28.09	14 <b>P</b> 30.97	15 <b>S</b> 32.06	16 <b>Cl</b> 35.45	17 <b>Ar</b> 39.95											18 <b>Kr</b> 83.80								
19 <b>K</b> 39.10	20 <b>Ca</b> 40.08	21 <b>Sc</b> 44.96	22 <b>Ti</b> 47.87	23 <b>V</b> 50.94	24 <b>Cr</b> 52.00	25 <b>Mn</b> 54.94	26 <b>Fe</b> 55.85	27 <b>Co</b> 58.93	28 <b>Ni</b> 58.69	29 <b>Cu</b> 63.55	30 <b>Zn</b> 65.38	31 <b>Ga</b> 69.72	32 <b>Ge</b> 72.63	33 <b>As</b> 74.92	34 <b>Se</b> 78.97	35 <b>Br</b> 79.90	36 <b>Kr</b> 83.80									
37 <b>Rb</b> 85.47	38 <b>Sr</b> 87.62	39 <b>Y</b> 88.91	40 <b>Zr</b> 91.22	41 <b>Nb</b> 92.91	42 <b>Mo</b> 95.95	43 <b>Tc</b> [97]	44 <b>Ru</b> 101.07	45 <b>Rh</b> 102.91	46 <b>Pd</b> 106.42	47 <b>Ag</b> 107.87	48 <b>Cd</b> 112.41	49 <b>In</b> 114.82	50 <b>Sn</b> 118.71	51 <b>Sb</b> 121.76	52 <b>Te</b> 127.60	53 <b>I</b> 126.90	54 <b>Xe</b> 131.29									
37 <b>Cs</b> 132.91	56 <b>Ba</b> 137.33											72 <b>Hf</b> 178.49	73 <b>Ta</b> 180.95	74 <b>W</b> 183.84	75 <b>Re</b> 186.21	76 <b>Os</b> 190.23	77 <b>Ir</b> 192.22	78 <b>Pt</b> 195.08	79 <b>Au</b> 196.97	80 <b>Hg</b> 200.59	81 <b>Tl</b> 204.38	82 <b>Pb</b> 207.2	83 <b>Bi</b> 208.98	84 <b>Po</b> [209]	85 <b>At</b> [210]	86 <b>Rn</b> [222]
87 <b>Fr</b> [223]	88 <b>Ra</b> [226]											104 <b>Rf</b> [267]	105 <b>Db</b> [268]	106 <b>Sg</b> [269]	107 <b>Bh</b> [270]	108 <b>Hs</b> [269]	109 <b>Mt</b> [277]	110 <b>Ds</b> [281]	111 <b>Rg</b> [282]	112 <b>Cn</b> [285]	113 <b>Nh</b> [286]	114 <b>Fl</b> [290]	115 <b>Mc</b> [290]	116 <b>Lv</b> [293]	117 <b>Ts</b> [294]	118 <b>Og</b> [294]
57 <b>La</b> 138.91	58 <b>Ce</b> 140.12	59 <b>Pr</b> 140.91	60 <b>Nd</b> 144.24	61 <b>Pm</b> [145]	62 <b>Sm</b> 150.36	63 <b>Eu</b> 151.96	64 <b>Gd</b> 157.25	65 <b>Tb</b> 158.93	66 <b>Dy</b> 162.50	67 <b>Ho</b> 164.93	68 <b>Er</b> 167.26	69 <b>Tm</b> 168.93	70 <b>Yb</b> 173.05	71 <b>Lu</b> 174.97												
89 <b>Ac</b> [227]	90 <b>Th</b> 232.04	91 <b>Pa</b> 231.04	92 <b>U</b> 238.03	93 <b>Np</b> [237]	94 <b>Pu</b> [244]	95 <b>Am</b> [243]	96 <b>Cm</b> [247]	97 <b>Bk</b> [247]	98 <b>Cf</b> [251]	99 <b>Es</b> [252]	100 <b>Fm</b> [257]	101 <b>Md</b> [258]	102 <b>No</b> [259]	103 <b>Lr</b> [262]												