



# ScienceGuyz

## CHEM 1212 Mock Exam 2 **KEY**

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## Spring 2020

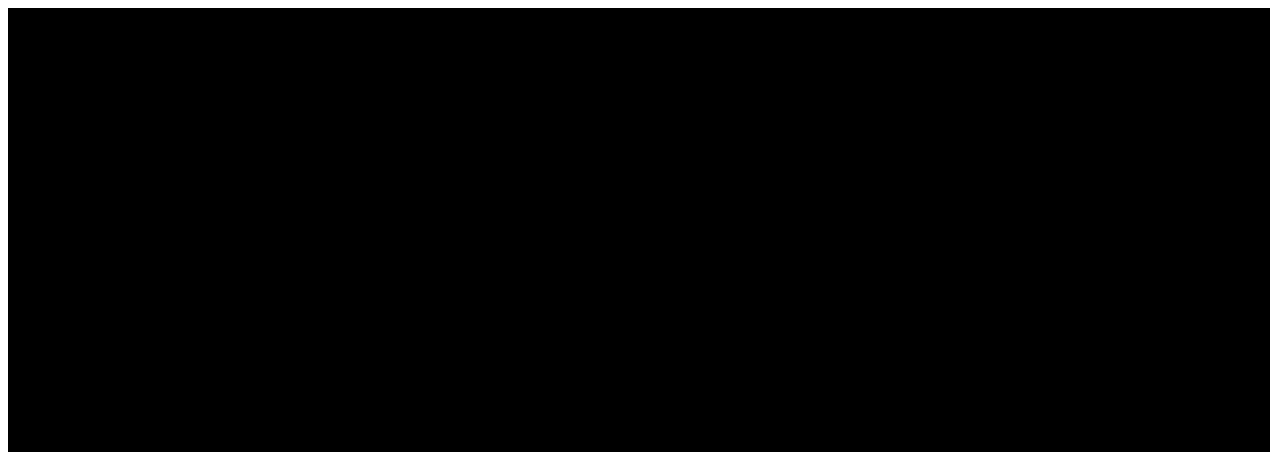
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### Multiple Choice

Identify the choice that best completes the statement or answers the question.

1. [REDACTED]  
[REDACTED]  
[REDACTED]



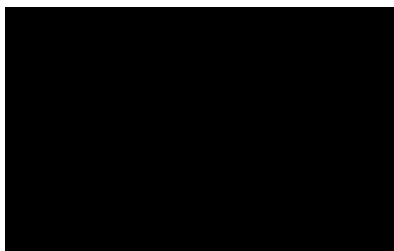
[REDACTED]  
[REDACTED]  
c) 2<sup>nd</sup> order.

[REDACTED]  
[REDACTED]

Be comfortable with how to determine rate law based on graphs. If you are given multiple graphs and asked to determine rate law, the graph with the straightest slope will represent the rate law. Once you determine the graph, look at the units provided on the y-axis to determine the rate law. 1/conc is indicative of a 2<sup>nd</sup> order reaction.

2. [REDACTED]  
[REDACTED]  
[REDACTED]  
[REDACTED]

b)  $\frac{-1}{2} \times \frac{\Delta[A]}{\Delta t} = \frac{1}{3} \times \frac{\Delta[D]}{\Delta t}$



Based on coefficients present, we know that the rate will be ½ with respect to [A] and 1/3 with respect to B. Because [A] is related to rate of disappearance, we must put a negative sign in front of the ½.

3.

[REDACTED]

[REDACTED]

[REDACTED]

[REDACTED]

[REDACTED]  
b)  $k[A]^2[B]$   
[REDACTED]  
[REDACTED]  
[REDACTED]

$$\text{rate} = k[A]^x[B]^y$$

Looking at experiment 2: when you quadruple the concentration of B, the rate of the reaction increases by quadruple as well (approximately), which indicates that the reaction is first order with [B].

Looking at experiment 3: when you double the concentration of A, the rate increases exponentially, which indicates that the reaction is second order in terms of [A].

This means that the rate for the reaction is  $\text{rate} = k[A]^2[B]$

4.

[REDACTED]  
[REDACTED]  
[REDACTED]  
[REDACTED]

[REDACTED]  
[REDACTED]  
[REDACTED]  
[REDACTED]

d) The reaction rate will decrease by a factor of 4.

[REDACTED]

You need to be very comfortable with the language of rate laws to answer a question like this. You are told that the reaction has two reactants and is third order. You are also told that when you double the concentration of the first reactant, the rate increases by a factor of 2. This indicates of the reaction being first order with respect to the first reactant. Because the overall reaction is third order and we know that it is first order with respect to reactant 1, we know it must be second order with respect to reactant 2. If you cut the concentration of reactant 2 in half, the reaction rate will decrease by a factor of 4.  $(1/2)^2$ .

5. [REDACTED]

a) At the same temperature, the rate of reaction B is greater than the rate of reaction A.

- ☐ [REDACTED]
- ☐ [REDACTED]
- ☐ [REDACTED]
- ☐ [REDACTED]
- ☐ [REDACTED]

Lower activation energy means that a reaction is more likely to proceed. Since reaction A has a higher activation energy, it will have slower rate than reaction B at the same temperature.

6. [REDACTED]

- ☐ [REDACTED]
- ☐ [REDACTED]
- ☐ [REDACTED]
- ☐ [REDACTED]

e) B and C are correct.

Activation energy increasing as temperatures go up would make the reaction go slower, not faster. In addition, you need a catalyst to affect activation energy. Activation energy decreasing would result in the reaction proceeding more quickly. Increasing temperature can cause the rate constant to increase, which results in the reaction occurring faster. Collisions increasing as temperatures increase is consistent with the collision theory, which also causes the reaction to proceed more quickly.

7. [REDACTED]

a) 6.21 seconds.

- ☐ [REDACTED]
- ☐ [REDACTED]
- ☐ [REDACTED]
- ☐ [REDACTED]

We are going to use the equation  $[A]_0 - [A]_t = -kt$

We are given that the  $[A]_0 = 0.612 \text{ M}$

$[A]_t = 0.115 \text{ M}$

$k = 0.080 \text{ M s}^{-1}$

Plug these numbers into the equation and solve!

$[A]_0 - [A]_t = -kt \rightarrow 0.612 \text{ M} - 0.115 \text{ M} = -(0.080 \text{ M s}^{-1}) t$

$t = 6.21 \text{ s}$

8.

- ☐ [redacted]
- ☐ [redacted]
- ☐ [redacted]
- d) 35.5 seconds.
- ☐ [redacted]

$$\ln\left(\frac{A}{A_0}\right) = -akt \rightarrow \ln\left(\frac{40}{100}\right) = -(1)(k)(47 \text{ sec})$$

$$k = 0.0195 \text{ sec}^{-1} \quad t_{1/2} = 0.693/0.0195 = \underline{35.5 \text{ sec}}$$

9.

- ☐ [redacted]
- ☐ [redacted]
- ☐ [redacted]
- d) 7
- ☐ [redacted]

$$0.188 \text{ g remain} = \frac{24 \text{ g initial}}{2^n}$$

$$n = \text{half life} \rightarrow 2^n = \frac{24}{0.188} = 128 \rightarrow n \ln(2) = \ln(128) \rightarrow n = 7$$

10.

- ☐ [redacted]
- ☐ [redacted]
- ☐ [redacted]
- b) Substance B is a catalyst; substance C is an intermediate.
- ☐ [redacted]
- ☐ [redacted]
- ☐ [redacted]

You may be given a graph to answer a question like this, or you may just be given text information. Understand a catalyst is going to be regenerated at some point in a reaction. We are told that substance B is a reactant and then later regenerated, so we know this to be a catalyst. We are told that substance C shows up, is consumed, and then is never seen again. This is indicative of an intermediate.

11. [REDACTED]

[REDACTED]

[REDACTED]

[REDACTED]

c)  $\text{Rate} = k[A]^2[B]^3$

[REDACTED]

[REDACTED]

Remember that the slowest step is the rate determining step and that with this, we only look at the reactants of the slowest step to determine the rate. This means that the rate is:

$$\text{rate} = k[A]^2[B]^3$$

12. [REDACTED]

[REDACTED]

[REDACTED]

[REDACTED]

[REDACTED]

[REDACTED]

e) Need additional information to answer question.

We are not given enough information in this problem to make any conclusions. If we have positive values for both enthalpy and entropy, we need temperature information to make any conclusions about the state of the reaction.

13. [REDACTED]

[REDACTED]

[REDACTED]

[REDACTED]

b)  $\Delta G = 0$

[REDACTED]

[REDACTED]

[REDACTED]

If you are told that the reaction is in a state of equilibrium, then the value of  $\Delta G$  must be equal to 0.

14. [REDACTED]

- ☐ [REDACTED]
- ☐ [REDACTED]
- ☐ [REDACTED]

[REDACTED]

[REDACTED]

☐ [REDACTED]

☐ [REDACTED]

☐ [REDACTED]

d) I and III

☐ [REDACTED]

From solid  $\rightarrow$  liquid  $\rightarrow$  gas entropy increases.

From gas  $\rightarrow$  liquid  $\rightarrow$  solid entropy decreases.

With this information, options I and III would have negative entropy values.

15. [REDACTED]

[REDACTED]

[REDACTED]

a)  $\Delta S$  must be a positive value,  $\Delta H$  must be a positive value.

☐ [REDACTED]

☐ [REDACTED]

☐ [REDACTED]

☐ [REDACTED]

You must be comfortable with working with the equation  $\Delta G = \Delta H - T\Delta S$ ! If you are told that the reaction is only spontaneous at high temperatures, that means that the values for  $\Delta H$  and  $\Delta S$  must both be positive.

16. [REDACTED]

☐ [REDACTED]

☐ [REDACTED]

☐ [REDACTED]

☐ [REDACTED]

☐ [REDACTED]

☐ [REDACTED]

d) I and II.

☐ [REDACTED]

If given thermodynamic data, you would be able to determine the equilibrium constant and the direction the spontaneous reaction would occur but would tell you nothing in terms of speed of the reaction. Remember from your study of kinetics that you need to perform an experiment to determine the kinetics of a reaction.

17. [REDACTED]
- [REDACTED]
- [REDACTED]
- [REDACTED]
- [REDACTED]
- [REDACTED]

a) I only.

- [REDACTED]
- [REDACTED]
- [REDACTED]
- [REDACTED]

From solid  $\rightarrow$  liquid  $\rightarrow$  gas: entropy increases.

From gas  $\rightarrow$  liquid  $\rightarrow$  solid: entropy decreases.

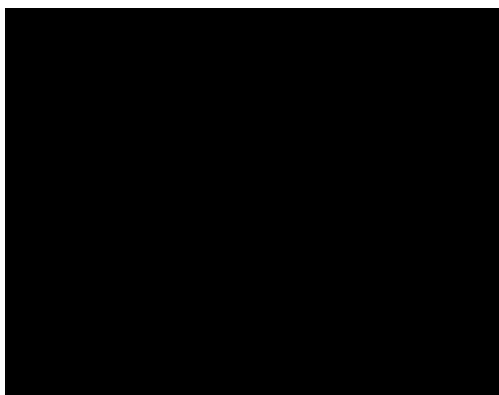
I. Freezing water = liquid  $\rightarrow$  solid, which mean that the entropy decreases.

II. Melting ice = solid  $\rightarrow$  liquid, which means that the entropy increases.

III. Sublimation = solid  $\rightarrow$  gas, which means that the entropy increases.

IV. Dissolving NaCl in water will cause entropy to increase.

18. [REDACTED]
- [REDACTED]
- [REDACTED]



- [REDACTED]
- [REDACTED]
- c) 178.8 J/mol . k
- [REDACTED]
- [REDACTED]

$$\Delta S = [4 (210.8) + 6 (188.8)] - [4 (192.8) + 5 (205.2)] = 178.8 \text{ J/mol} \cdot \text{k}$$



19. [REDACTED]  
[REDACTED]  
[REDACTED]  
[REDACTED]  
b) -348 J/K.  
[REDACTED]  
[REDACTED]  
[REDACTED]

$$\Delta H_{\text{vap}} = 30.67 \text{ kJ/mol}$$

To determine amount of energy, multiply  $\Delta H_{\text{vap}}$  by the number of moles you have present.

$$\frac{120 \text{ g CH}_3\text{OH}}{32 \text{ g}} \times 1 \text{ mol} = 3.75 \text{ mol}$$

$$\frac{30.67 \text{ kJ}}{1 \text{ mol}} \times 3.75 \text{ mol} = 115.013 \text{ kJ} = 115013 \text{ J}$$

Condensation is an exothermic process; therefore, we need to put a negative sign in front of the value we obtained: 115013 J = -115013 J


Make sure to convert temperature to kelvin!  $57.5^\circ\text{C} + 273.15 = 330.65 \text{ K}$



$$S = \frac{-115013 \text{ J}}{330.65 \text{ K}} = -348 \text{ J/K}$$


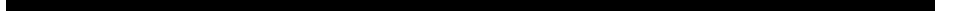
20. [REDACTED]  
[REDACTED]  
[REDACTED]  
[REDACTED]  
[REDACTED]  
d) Reversible reactions can be reversed with no net change in either system or surroundings.  
[REDACTED]



Reversible reactions are spontaneous in one direction and nonspontaneous in another, meaning option A is incorrect. Reversible reactions do not necessarily have to be carried out at a specific condition, meaning options B and C are incorrect.

For reversible reactions, one direction is endergonic (a positive  $\Delta G$  value) and the other is exergonic (a negative  $\Delta G$  value). If a reversible reaction proceeds in one direction, it can be reversed with no net change in energy of the system. For example, if a reaction is exergonic (spontaneous) in one direction, the amount of energy released can be used to drive the endergonic (non-spontaneous) in the other direction, leaving no net change in energy. This makes option D correct.



21. 

d) The concentration of each substance in the vessel will be constant.

Option A is incorrect as the rate of the forward and reverse reactions are identical, but not the moles of products and reactants. Option B is incorrect as further reaction can occur. Equilibrium does not mean that the reaction stops happening. As mentioned in the previous question, a catalyst will have no impact on the equilibrium position, making option C incorrect. Option D is the only correct statement regarding equilibrium, as the concentration of each substance in the vessel will be constant

22. [REDACTED]  
[REDACTED]  
[REDACTED]

a)  $4.41 \times 10^{-4}$

□	□
□	□
□	□
□	□

$$K_c = \frac{\text{products}}{\text{reactants}} \rightarrow \frac{[0.00075]^2}{[0.025][0.051]} = 4.41 \times 10^{-4}$$

23. [REDACTED]

**a)**

**b)**

c)  $K_c$  is less than  $Q$ ; more HI will be made.

$$Q = \frac{[.36][.36]}{[1.5]^2} = 0.0576$$

Based on this, we know that  $Q > K_c$ . Since  $Q$  is greater than  $K_c$  ( $K_c$  is less than  $Q$ ), we know that the reaction will shift to the reactants and more HI will be produced.

24. [REDACTED]

[REDACTED]

[REDACTED]  
[REDACTED]  
[REDACTED]  
[REDACTED]

e)  $K_p = K_c(RT)^{-1}$

You know that the relationship between  $K_c$  and  $K_p$  is given by the equation  $K_p = K_c(RT)^{\Delta n}$   
In this example,  $\Delta n = 2 - 3 = -1$

Therefore,  $K_p = K_c(RT)^{-1}$

25. [REDACTED]

[REDACTED] [REDACTED] [REDACTED] [REDACTED] [REDACTED] [REDACTED]  
[REDACTED] [REDACTED] [REDACTED] [REDACTED] [REDACTED] [REDACTED]

[REDACTED]

[REDACTED] [REDACTED] [REDACTED] [REDACTED]

a) 5.377.

[REDACTED]  
[REDACTED]  
[REDACTED]  
[REDACTED]

Flip & x 4 for reaction 2!

Therefore  $\rightarrow \frac{1}{(0.771)^4}$  for reaction 2

$K_c \text{ overall} = (1.90) \left( \frac{1}{(0.771)^4} \right) = 5.38$

26. [REDACTED]

[REDACTED]  
[REDACTED]  
[REDACTED]  
[REDACTED]

[REDACTED]  
[REDACTED]  
[REDACTED]  
[REDACTED]

e) I, III, and IV.

Understand that changing the heat levels, changing the concentration of reactant and changing the volumes of reactants will all have an impact on the equilibrium position. Of the options given, the only one that has no effect is a catalyst. Adding or removing a catalyst will have no impact on the equilibrium position.

27. [REDACTED]

[REDACTED]

[REDACTED]

[REDACTED]

☐ [REDACTED]

b) The partial pressure of  $\text{CO}_2$  in the reaction vessel has not changed.

☐ [REDACTED]

☐ [REDACTED]

☐ [REDACTED]

$\text{CaCO}_3$  is a solid and thus not considered when looking at equilibrium expression. Because we do not consider this compound in the equilibrium expression, adding more will have no impact. The partial pressure of  $\text{CO}_2$  remains unchanged.

28. [REDACTED]

[REDACTED]

[REDACTED]

☐ [REDACTED]

☐ [REDACTED]

☐ [REDACTED]

d) The volume is increased.

☐ [REDACTED]

Shift to the left means a shift towards the reactants. There are more moles of gas on the reactant side than product side. Remember that an increase in pressure means a decrease in volume. Increasing the pressure will cause the reaction to shift to the side with less moles of gas (right). Increasing the volume will cause the reaction to shift to left (more moles of gas).

29. [REDACTED]

[REDACTED]

[REDACTED]  
 [REDACTED]  
 [REDACTED]  
 [REDACTED]  
 [REDACTED]  
 [REDACTED]  
 [REDACTED]  
 [REDACTED]

d) The rate of the forward reaction will increase while the rate of the reverse reaction decreases.

[REDACTED]  
 [REDACTED]

Remember that if the reaction is exothermic, the heat will show up on the product side. Decreasing temperature will cause the reaction to shift to the right, which means an increase in the forward reaction.

30.

[REDACTED]

3D (g)

[REDACTED]  
 [REDACTED]  
 [REDACTED]  
 [REDACTED]

c) 1) shift towards products; 2) shift towards products; 3) shift toward reactants; 4) shift towards products.

[REDACTED]  
 [REDACTED]  
 [REDACTED]  
 [REDACTED]

Increase volume = shift toward the side with the more moles of gas (products)  
 Increase temperature = shift toward products because the reaction is endothermic  
 [C][D] increase = shift toward reactants  
 [A][B] increase = shift toward products

31.

e) 1.01 g CS<sub>2</sub> remaining; 2.10 g CS formed.

$$\frac{63 \text{ days}}{1 \text{ day}} \times \frac{24 \text{ hours}}{1 \text{ day}} \times \frac{3600 \text{ s}}{1 \text{ hr}} = 5.44 \times 10^6 \text{ seconds.}$$

$$\ln \left( \frac{A}{4.64} \right) = -(1)(2.8 \times 10^{-7})(5.44 \times 10^6 \text{ seconds}) = \underline{1.01 \text{ g CS}_2 \text{ remaining}}$$

$$4.64 \text{ g CS}_2 \text{ initial} - 1.01 \text{ g CS}_2 \text{ remaining} = 3.63 \text{ g CS}_2 \text{ reacted}$$

$$\frac{3.63 \text{ g CS}_2}{76 \text{ g CS}_2} \times \frac{1 \text{ mol CS}_2}{1 \text{ mol CS}_2} \times \frac{1 \text{ mol CS}}{1 \text{ mol CS}_2} \times \frac{44 \text{ g CS}}{1 \text{ mol CS}} = \underline{2.10 \text{ g CS formed}}$$

32.

d) -79.406 kJ/mol.

$$\Delta G = \Delta G^\circ + R \ln Q$$

$$\Delta G = -32.7 \frac{\text{kJ}}{\text{mol}} + [0.008314(298.15) \ln \left( \frac{[1.2 \times 10^{-9}]^2}{[0.0026][0.0044]^3} \right)] = -79.406 \text{ kJ/mol}$$