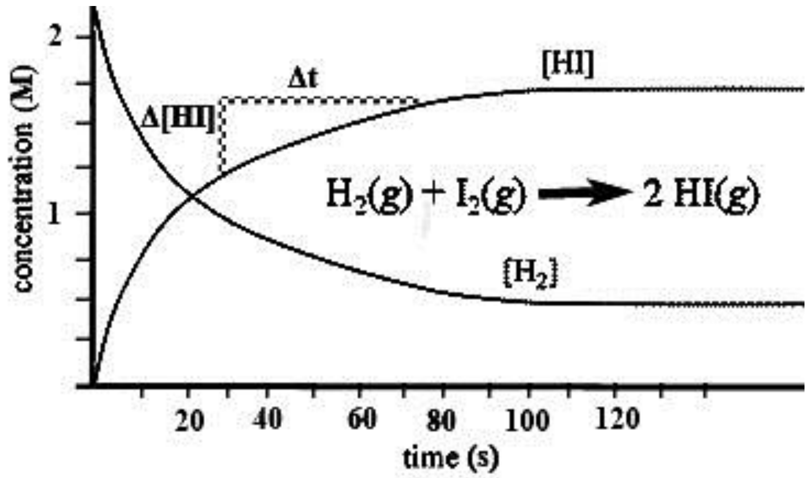
1. Given the graphed data at right, if you are asked to calculate the concentration change for HI(g) during the time interval from 30 s to 75 s, you would be calculating the



A. rate law.

B. rate constant.

C. instantaneous rate.

D. average rate.

2. For the reaction A + B → 2 C, use the table below to determine $\frac{\Delta[B]}{\Delta t}$.

Time (s)	[A] (M)	[B] (M)	[C] (M)
10.0	0.910	0.500	0.250
20.0	0.730	0.320	0.610

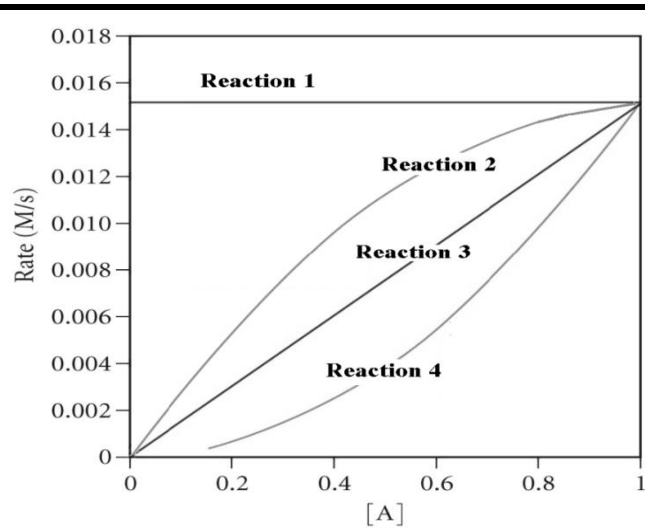
A. 1.80×10^{-2} M/s

B. 9.00×10^{-3} M/s

C. -1.80×10^{-2} M/s

D. -3.60×10^{-2} M/s

3. According to the plot, which of these reactions demonstrates zero-order kinetics?



A. Reaction 1

B. Reaction 2

C. Reaction 3

D. Reaction 4

4. For the reaction 3 AB → 2 C, use the table below to determine the **order** of reaction:

[AB] (M)	Initial Rate (M/s)
0.046	0.0042
0.023	0.0011

A. 0
B. 1
C. 2
D. 3

5. For the reaction

$$H_2SeO_3(aq) + 6\ I^-(aq) + 4\ H^+(aq) \rightarrow Se(s) + 2\ I_3^-(aq) + 3\ H_2O(l)$$

$$\text{rate} = k[H_2SeO_3]^a[H^+]^b[I^-]^c$$

it is experimentally determined that the reaction rate doubles when $[H_2SeO_3]$ is doubled, the rate quadruples when $[H^+]$ is doubled, and the rate increases by a factor of eight when $[I^-]$ is doubled. What is the **overall order** of the reaction?

A. third order
B. fourth order
C. sixth order
D. eighth order

6. The decomposition of $N_2O_5(g)$ is a first-order reaction with $k = 7.0 \times 10^{-2}\ s^{-1}$. If the initial concentration of $N_2O_5(g)$ is 1.5 M, what is the concentration of N_2O_5 after 4.2 minutes?

A. 7.8 M
B. 4.2×10^{-4} M
C. 3.3×10^{-8} M
D. 2.1×10^{-9} M

7. The reaction of butadiene, C_4H_6, to form C_8H_{12} is second order in butadiene.

$$2\ C_4H_6(g) \rightarrow C_8H_{12}(g) \qquad k = 6.14 \times 10^{-2}\ M^{-1}s^{-1}.$$

<u>How long</u> has the reaction run if the initial butadiene concentration was 0.278 M and the current butadiene concentration is 0.074M?

A. 18 minutes

B. 11 minutes

C. 2.7 minutes

D. 1.3 minutes

8. The decomposition of A is a first-order reaction with a half-life of 738 seconds. The initial concentration of A is 0.503 M. Determine the rate constant, *k*.

$$A \rightarrow \text{products}$$

A. $1.04 \times 10^{3}\ s^{-1}$

B. $5.36 \times 10^{-2}\ s^{-1}$

C. $1.94 \times 10^{-3}\ s^{-1}$

D. $9.39 \times 10^{-4}\ s^{-1}$

9. What effect does doubling the Kelvin temperature of a reaction mixture have on the rate constant, *k*?

A. The value of the rate constant doubles.

B. The value of the rate constant increases, but not necessarily doubles.

C. The value of the rate constant decreases by half.

D. The value of the rate constant decreases, but not necessarily by half.

10. According to the collision model for a chemical reaction, the frequency factor A in the Arrhenius equation depends on the

A. orientation factor and collision frequency.

B. temperature and concentration.

C. activation energy and temperature.

D. activation energy and orientation factor.

11. The rate of the overall reaction

$$2\ H_2O_2(aq) \rightarrow 2\ H_2O(l) + 2\ O_2(g)$$

can be increased by adding bromine. According to these elementary steps, what are the functions of Br_2, Br^-, and H^+ in the overall reaction?

$$Br_2(aq) + H_2O_2(aq) \rightarrow 2\ Br^-(aq) + 2\ H^+(aq) + O_2(g)$$
$$2\ Br^-(aq) + 2\ H^+(aq) + H_2O_2(aq) \rightarrow 2\ H_2O(l) + O_2(g) + Br_2(aq)$$

A. Br_2, Br^-, and H^+ are all catalysts.

B. Br^- and H^+ are catalysts; Br_2 is a reaction intermediate.

C. Br_2 is a catalyst; Br^- and H^+ are reaction intermediates.

D. Br_2, Br^-, and H^+ are all reaction intermediates.

12. For the reaction

$$H_2(g) + 2\ IBr(g) \rightarrow 2\ HBr(g) + I_2(g),$$

the following mechanism has been proposed.

Step 1: $H_2(g) + IBr(g) \rightarrow HBr(g) + HI(g)$
Step 2: $HI(g) + IBr(g) \rightarrow HBr(g) + I_2(g)$

If the rate law is rate = $k[H_2][IBr]$, which is the rate-determining step in the proposed mechanism?

A. Step 1 is the rate-determining reaction.

B. Step 1 and Step 2 both determine the rate.

C. Step 2 is the rate-determining step.

D. The rate-determining step cannot be identified from the information given.

13. Catalysts

A. increase the rate of reaction by lowering the activation energy, E_a.

B. increase the rate of the reaction by increasing the reaction temperature, T.

C. increase the rate of reaction by adding energy to the reaction.

D. must be in the same phase as the reactants and products.

14. At dynamic equilibrium,

A. the concentrations of the reactants and products are equal.

B. the concentrations of the reactants and products remain constant.

C. the rates of the forward and reverse reactions are equal.

D. both B and C are correct.

15. Initially, 0.800 M of SO_2 and 0.200 M of O_2 are present in a reaction vessel. Given the balanced chemical equation and K_c, what is true about the **equilibrium** concentrations of reactants and products?

$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g) \qquad K_c = 1.98 \text{ at } 1303\text{ K}$$

A. The equilibrium strongly favors neither reactants nor products, so appreciable amounts of SO_3, SO_2 and O_2 will all be present at equilibrium.

B. The equilibrium lies far to the right and the concentration of SO_3 will be significantly higher than the concentrations of SO_2 and O_2.

C. The equilibrium lies far to the right and the concentration of SO_3 will be significantly lower than the concentrations of SO_2 and O_2.

D. The equilibrium lies far to the left and the concentration of SO_3 will be significantly lower than the concentrations of SO_2 and O_2.

16. Given the reactions

$$2\ NO(g) \rightleftharpoons N_2(g) + O_2(g) \qquad K_{c1} = 4.4 \times 10^{18}$$
$$2\ NO(g) + O_2(g) \rightleftharpoons 2\ NO_2(g) \qquad K_{c2} = 3.0 \times 10^{6}$$

calculate K_{c3} for the reaction

$$N_2(g) + 2\ O_2(g) \rightleftharpoons 2\ NO_2(g) \qquad K_{c3} = ?$$

A. 6.8×10^{-13}

B. 2.2×10^{-6}

C. 3.1×10^{11}

D. 8.4×10^{24}

17. Which of the following is the equilibrium constant expression for this reaction?

$$FeO(s) + H_2(g) \rightleftharpoons Fe(s) + H_2O(g)$$

A. $K_c = \dfrac{[H_2O]^2}{[FeO_2][H_2]}$

B. $K_c = \dfrac{[Fe][H_2O]}{[FeO][H_2]}$

C. $K_c = \dfrac{[FeO][H_2]}{[Fe][H_2O]}$

D. $K_c = \dfrac{[H_2O]}{[H_2]}$

18. $N_2O_4(g)$ and $NO_2(g)$ are in equilibrium according to this equation.

$$N_2O_4(g) \rightleftharpoons 2\ NO_2(g) \qquad K_c = ?$$

$N_2O_4(g)$ at an initial concentration of 0.100 M produces a 0.132 M equilibrium concentration of $NO_2(g)$ at 500. K. Determine K_c at 500. K.

A. 0.0972

B. 0.210

C. 0.314

D. 0.512

19. The reaction

$$2\ NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$$

has an equilibrium constant, $K_P = 1.8 \times 10^{-2}$ at 800 °C. If the reaction quotient, $Q = 9.3 \times 10^{-4}$ at 800 °C,

A. the reaction will proceed towards the reactants (to the left).

B. the reaction has reached equilibrium and no change in concentration occurs.

C. the reaction will proceed towards the products (to the right).

D. more reactants are needed for the reaction to continue.

20. If the equilibrium concentrations of PCl_3 and Cl_2 are both 0.0852 M, what is the **equilibrium concentration** of PCl_5?

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g) \qquad K_c = 0.030$$

A. 0.013 M

B. 0.24 M

C. 0.76 M

D. 3.6 M

21. A reaction vessel is filled with 0.400 M of $N_2(g)$ and 0.200 M of $O_2(g)$. What is the **equilibrium concentration** of N_2O at 350 K?

$$2\ N_2(g) + O_2(g) \rightleftharpoons 2\ N_2O(g) \qquad K_c = 7.3 \times 10^{-36} \text{ at } 350 \text{ K}$$

A. 2.3×10^{-30} M

B. 4.8×10^{-19} M

C. 3.7×10^{-9} M

D. 1.6×10^{-5} M

22. Given the reaction,

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g) \qquad \Delta H = -41 \text{ kJ};\ K_c = 9.03 \text{ at } 698 \text{ K},$$

which change will cause the greatest shift in the equilibrium toward **products**?

A. decreasing the pressure

B. lowering the temperature

C. adding $Ar(g)$

D. adding $H_2(g)$

23. For the rection,

$$P_4(g) + 6\ Cl_2(g) \rightleftharpoons 4\ PCl_3(g)$$

at equilibrium at a fixed temperature, which of the following is **correct**?

A. Decreasing the volume causes the reaction to shift to the left (reactants).

B. Decreasing the volume causes the reaction to shift to the right (products).

C. Increasing the volume causes the reaction to shift to the right (products).

D. Increasing the volume has no effect on the composition of the equilibrium system..

24. Which acid-base theory describes the reactions below? What are the functions of the reactants?

Reaction 1: $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
Reaction 2: $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

A. Reaction 1 is best described by the Arrhenius theory; NaOH is the acid and HCl is the base. Reaction 2 is best described by the Brønsted-Lowry theory; NH_3 is the acid and HCl is the base.

B. Reaction 1 is best described by the Arrhenius theory; NaOH is the base and HCl is the acid. Reaction 2 is best described by the Brønsted-Lowry theory; NH_3 is the base and HCl is the acid.

C. Reaction 1 is best described by the Brønsted-Lowry theory; NaOH is the acid and HCl is the base. Reaction 2 is best described by the Arrhenius theory; NH_3 is the acid and HCl is the base.

D. Reaction 1 is best described by the Brønsted-Lowry theory; NaOH is the base and HCl is the acid. Reaction 2 is best described by the Arrhenius theory; NH_3 is the base and HCl is the acid.

25. Considering the acid equilibrium reaction

$$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq),$$

A. the weaker the acid, the weaker its conjugate base.

B. the stronger the acid, the weaker its conjugate base.

C. the stronger the acid, the stronger its conjugate base.

D. a large acid ionization constant indicates a very weak acid.

26. Calculate the pH of pure water **at 50 °C**. $K_w = 5.48 \times 10^{-14}$ at 50 °C.

A. 6.631
B. 6.916
C. 7.091
D. 7.362

27. Which statement is true about the following acid ionization reactions?

$$HNO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$$

$$HCHO_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CHO_2^-(aq)$$

A. HNO_3 is a strong acid and $[H_3O^+]$ is not equal to $[HNO_3]$.
B. HNO_3 is a weak acid and $[H_3O^+]$ is equal to $[HNO_3]$.
C. $HCHO_2$ is a strong acid and $[H_3O^+]$ is not equal to $[HCHO_2]$.
D. $HCHO_2$ is a weak acid and $[H_3O^+]$ is not equal to $[HCHO_2]$.

28. Which 0.100 M acid solution will have the **highest** percent ionization?

hypochlorous acid, $HClO$	$K_a = 2.9 \times 10^{-8}$
acetic acid, $HC_2H_3O_2$	$K_a = 1.8 \times 10^{-5}$
lactic acid, $HC_3H_5O_3$	$K_a = 1.4 \times 10^{-4}$
iodic acid, HIO_3	$K_a = 1.7 \times 10^{-1}$

A. hypochlorous acid
B. acetic acid
C. lactic acid
D. iodic acid

29. Determine the pH of a mixture of weak acids that is 0.200 M in HNO_2 and 0.150 M in HCN.

$$HNO_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NO_2^-(aq) \qquad K_a = 4.6 \times 10^{-4}$$
$$HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq) \qquad K_a = 4.9 \times 10^{-10}$$
$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq) \qquad K_w = 1.0 \times 10^{-14}$$

A. 2.02
B. 3.24
C. 4.18
D. 7.00

30. Determine the pH of a 0.27 M pyridine (C_5H_5N) solution.

$$C_5H_5N(aq) + H_2O(l) \rightleftharpoons C_5H_5NH^+(aq) + OH^-(aq) \qquad K_b = 1.7 \times 10^{-9}$$

A. 4.66
B. 8.77
C. 9.33
D. 13.43

Answer Key:

1. D
2. C
3. A
4. C
5. C
6. C
7. C
8. D
9. B
10. A
11. C
12. A
13. A
14. D
15. A
16. A
17. D
18. D
19. C
20. B
21. B
22. B
23. B
24. B
25. B
26. A
27. D
28. D
29. A
30. C