Name: Answer Key Lab Section: ____________

Exam #3

Mark your answers on this exam, showing all work (including units) where appropriate. The Periodic Table is on the last page. Each answer is worth 3 points.

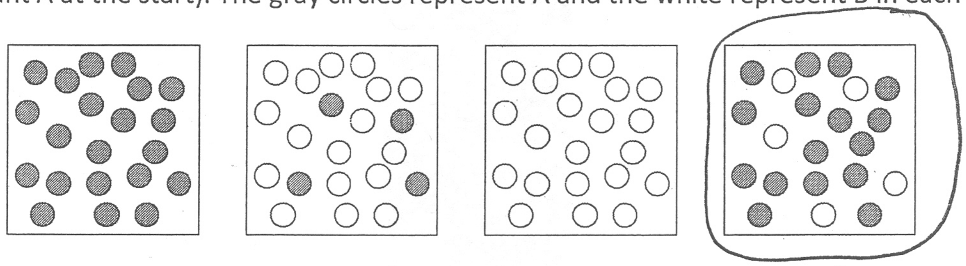
For a first order process: $\ln\frac{N_t}{N_o} = -kt \qquad N_t = N_o e^{-kt}$

1. When the reversible reaction, $N_2 + O_2 \leftrightarrow 2\ NO$, has reached a state of dynamic equilibrium, which statement(s) below is(are) true? (circle all true statements)

B

   a. Both the forward and reverse reactions shut down and no more NO, N2 or O2 are produced.
   b. The rate of the forward reaction equals the rate of the reverse reaction.
   c. The rate constant of the forward reaction equals the rate constant of the reverse reaction.

2. The equilibrium constant for the reaction A↔B is K_{eq}=0.0002 at room temperature. Circle the figure that best represents the system when it reaches equilibrium (assume there is 100% reactant A at the start). The gray circles represent A and the white represent B in each figure.



$K_{eq} = \frac{[B]}{[A]}$

If K_{eq} is small [A] > [B]

gray white

3. If the reaction A + B → C has the rate law: Rate = $k[A]^2[B]$, the units of the rate constant are:

   a. s^{-1}
   b. mol/L·s
   c. L/mol·s
   d. $L^2/mol^2 \cdot s$
   e. $L^3/mol^3 \cdot s$

D

$$\text{Rate}\left(\frac{mol}{L \cdot s}\right) = \underbrace{\frac{\cancel{L^2}}{\cancel{mol^2} \cdot s}}_{k} \underbrace{\left(\frac{\cancel{mol^2}}{\cancel{L^2}}\right)}_{[A^2]} \underbrace{\left(\frac{mol}{L}\right)}_{[B]}$$

4. For a reaction with the rate law: Rate = k[A]: ← First order equation

   a. a plot of ln[A] vs. time gives a straight line.
   b. a plot of 1/[A] vs. 1/time gives a straight line.
   c. a plot of [A] vs. time gives a straight line.
   d. a plot of 1/[A] vs. time gives a straight line.
   e. a plot of log $\frac{[A]_0}{[A]}$ vs. 1/time gives a straight line.

A

5. What is the osmotic pressure at 310 K of a sugar solution containing 42.0 g of sucrose ($C_{12}H_{22}O_{11}$) in 650 mL of solution?

$\Pi = CRT$

$C = \frac{42.0\,g \times \left(\frac{mol}{342\,g}\right)}{0.650\,L} = 0.189\,\frac{mol}{L}$

$\Pi = \left(\frac{0.189\,mol}{L}\right)\left(\frac{0.0821\,L \cdot atm}{K \cdot mol}\right)(310\,K)$

$\Pi = 4.8\,atm$

E

   a. 3.12 atm
   b. 0.065 atm
   c. 1.64 atm
   d. 1069 atm
   e. 4.80 atm

6. For the reaction:

$$BF_3\,(g) + NH_3\,(g) \rightarrow F_3BNH_3\,(g)$$

The rate law is determined to be: Rate = $k[BF_3][NH_3]$, where the value of $k = 3.41\ M^{-1}sec^{-1}$
What is the initial rate of this reaction when the concentration of the two reactants are $[BF_3]$ = 0.500 M and $[NH_3]$ = 0.500 M?

Rate = $(3.41\,M^{-1}sec^{-1})(0.5\,M)(0.5\,M)$

Rate = 0.852 M/sec

D

   a. 0.213 M/sec
   b. 0.25 M/sec
   c. 0.426 M/sec
   d. 0.852 M/sec
   e. 3.41 M/sec

7. In a study designed to prepare new gasoline-resistant coating, a polymer chemist dissolves 6.061 g poly(vinyl alcohol) in enough water to make 100.0 mL of solution. At 25°C, the osmotic pressure of this solution is 0.248 atm. What is the molar mass of the polymer sample?

$C = \frac{\Pi}{RT} = \frac{0.248\,atm}{\left(0.0821\,\frac{L\,atm}{K\,mol}\right)(298\,K)} = 0.0101\,\frac{mol}{L} \times 0.10\,L = 0.00101\,mol$

molar mass $= \frac{6.061\,g}{0.00101\,mol} = 5979\,g/mol = 5.98 \times 10^{3}\,\frac{g}{mol}$

D

   a. 0.0101 g/mol
   b. 73.2 g/mol
   c. 502 g/mol
   d. 5.98 x 10^3 g/mol
   e. 8.87 x 105 g/mol

8. Use the following information for the reaction: 2A + B → C. Initial rates were measured at different molar concentrations.

Experiment	[A]	[B]	Initial Rate (mol/L·s)
1	1.0 x 10^{-2}	1.0 x 10^{-2}	5.3 x 10^{-4}
2	2.0 x 10^{-2}	1.0 x 10^{-2}	1.06 x 10^{-3}
3	3.0 x 10^{-2}	2.0 x 10^{-2}	1.59 x 10^{-3}
4	2.0 x 10^{-2}	2.0 x 10^{-2}	1.06 x 10^{-3}

When B doubles, rate stays the same = zero order.

What is the order of this reaction with respect to reactant [B]?

A

   a. Zero
   b. One
   c. Two
   d. Three

B

9. What is a *half-life*?
   a. It is a special case when the rate constant equals one-half. (ie $k = 0.5$)
   b. A period of time required for the concentration of a reactant to reduce to one-half its original value.
   c. It is a rate law that is one-half order: (ie. Rate = $k[A]^{1/2}$)
   d. A period representing one-half the amount of time required for an initial concentration of a reactant to reduce to 0.00 M.
   e. A period representing the amount of time required for an initial concentration of a reactant to reduce to 0.5 M.

C

10. The gas phase reaction $C_2H_4 + Cl_2 \rightarrow C_2H_4Cl_2$ follows the rate law: Rate = $k[C_2H_4][Cl_2]^2$. If the concentration of C_2H_4 is doubled while the concentration of Cl_2 remains constant, the initial rate of the reaction:
    a. increases by a factor of 4.
    b. decreases by a factor of 2.
    c. increases by a factor of 2.
    d. increases by a factor of 1.5.
    e. remains constant.

Rate = $k[C_2H_4][Cl_2]^2$

If this kept constant, rate depends on order with respect to C_2H_4. The reaction is 1st order in C_2H_4, so if concentration is doubled, the rate will double.

B

11. The highly toxic pesticide parathion, $(C_2H_5O)_2PSOC_6H_4NO_2$, decomposes in the soil via a *first order rate law*. The rate constant for the decomposition of parathion is 0.0495 days^{-1}. How long will it take for 50% of the parathion to decompose after it is applied to the land?
    a. 1.1 week
    b. 2 weeks
    c. 3 weeks
    d. 4 weeks
    e. 5 weeks

$N_t = N_0 e^{-kt_{1/2}}$ $\quad t_{1/2} = \frac{0.693}{k} = \frac{0.693}{0.0495 \text{ days}^{-1}}$

$t_{1/2} = 14$ days

A or E

12. For the reaction $2\,SO_2(g) + Cl_2(g) \rightleftarrows 2\,SO_2Cl_2(g)$, the expression for K_{eq} is

a. $K_{eq} = \frac{[SO_2Cl_2]^2}{[SO_2]^2[Cl_2]^2}$

b. $K_{eq} = \frac{[SO_2]^2[Cl_2]}{[SO_2Cl_2]^2}$

c. $K_{eq} = \frac{2[SO_2Cl_2]^2}{2[SO_2]^2[Cl_2]}$

d. $K_{eq} = \frac{[SO_2Cl_2]}{[SO_2]^2[Cl_2]}$

e. $K_{eq} = \frac{[SO_2Cl_2]^2}{[SO_2]^2[Cl_2]}$

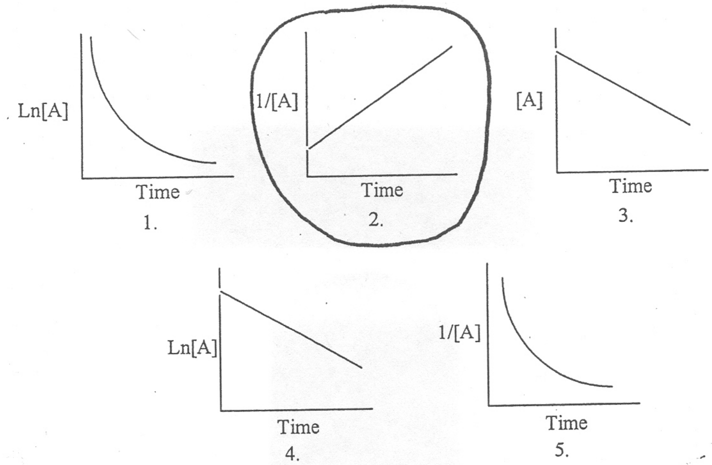
As written, the reaction is unbalanced. The K_{eq} for this reaction is e.
If the reaction is properly balanced, it looks like this:
$2SO_2(g) + 2Cl_2(g) \rightleftarrows 2SO_2Cl_2(g)$
and K_{eq} would be a.

A

13. Iodine-123 is used to study thyroid gland function. This radioactive isotope breaks down in a *first order* process with a half-life of 13.1 hours. What is the rate constant, k for the process?
    a. $k = 0.0529$ hr^{-1}
    b. $k = 6.55$ hr^{-1}
    c. $k = 9.08$ hr^{-1}
    d. $k = 18.9$ hr^{-1}
    e. $k = 26.2$ hr^{-1}

$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{13.1 \text{hr}} = 0.053/\text{hr}$

14. Graphical analysis can be used to determine the rate law for a general reaction of reactant "A" going to products "B" and "C". Circle the plot that accurately represents the data if this particular reaction follows a *second order rate law*?



15. The rate constant for the first order dehydration of tert-butyl alcohol at 773 K (T_1) is $1.20 \times 10^{-4}\ s^{-1}$. The rate constant for this process at 873 K (T_2) is $6.80 \times 10^{-3}\ s^{-1}$. The activation energy for this reaction in kJ/mol is: ($R = 8.314$ J/mol·K)

B

- a. 27.4
- b. 226
- c. 318
- d. 3.36
- e. 22

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

$R = 8.314\ J/K \cdot mol$

$$\ln\left(\frac{6.8 \times 10^{-3}}{1.2 \times 10^{-4}}\right) = \frac{-E_a}{R}\left(\frac{1}{873} - \frac{1}{773}\right)$$

$$56.66667 = \frac{-E_a}{R}(-1.48186 \times$$

16. The equilibrium constant for the reaction below is $K_{eq}(25°C) = 3.2 \times 10^{-34}$

$$2\ HCl\ (g) \rightleftarrows H_2\ (g) + Cl_2\ (g) \qquad \frac{[H_2][Cl_2]}{[HCl]^2}$$

What can be stated about the relative equilibrium distribution for this mixture?

C

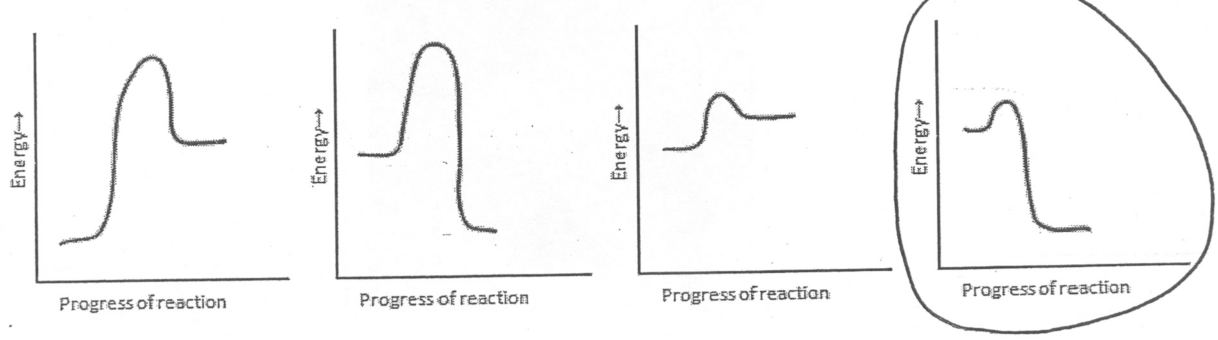
- a. The mixture contains predominantly H_2.
- b. The mixture contains predominantly Cl_2.
- c. The mixture contains predominantly HCl.
- d. The mixture contains predominantly H_2 and Cl_2.
- e. The mixture contains relatively equal amounts of H_2, Cl_2 and HCl.

Since Keq is very small, @ equilibrium the [HCl] is high

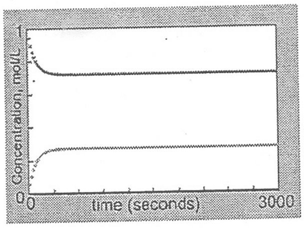
17. Which Reaction Energy Diagram shown below represents a reaction that proceeds *fast* and is exothermic?

→ Releases E (products lower than reactants)

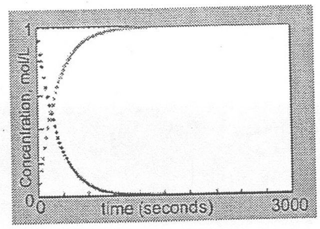
→ small Ea (bump)



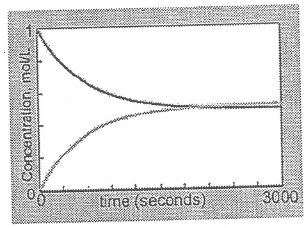
18. The following figures represent a reaction A↔B, where A is represented by the black line and B by the light gray line:



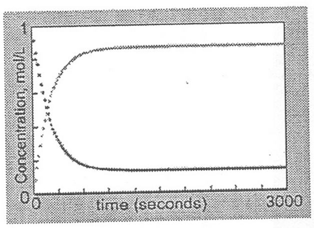
a.



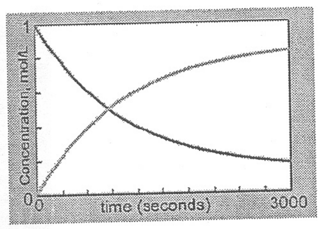
b.



c.



d.



e.

$K_{eq} = \frac{[B]}{[A]}$

B a. Which reaction goes to completion? b.

C b. Which reaction has an equilibrium constant that is close to 1? C. ($[B] \approx [A]$ @ equilibrium)

A c. Which reaction has the lowest equilibrium constant? a.
↳ least amount of B is made

C d. For which reaction are the equilibrium concentrations of A and B almost equal? C.

19. A reaction has the equilibrium expression

$$K = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

a. Write the chemical equation for this reaction.

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

b. Given the equilibrium concentrations in the table, what is the value of the equilibrium constant?

	Equilibrium Concentration
SO_2	0.0124
O_2	0.031
SO_3	0.037

$$K = \frac{[0.037]^2}{[0.0124]^2[0.031]}$$

$K = 287.2$

20. Consider the following two-step mechanism that has been proposed for the gas phase reaction of H_2 with ICl to produce I_2 and HCl :

Step 1: H_2 (g) + ICl (g) → ~~HI (g)~~ + HCl (g) (slow) Rate = k [H_2][ICl]

Step 2: ~~HI (g)~~ + ICl (g) → I_2 (g) + HCl (g) (fast)

Net reaction: H_2 (g) + 2 ICl (g) → I_2 (g) + 2 HCl

a. What is the *rate law* for the net overall reaction?

Rate = k[H_2][ICl]

b. What *intermediate*(s) is(are) present in this mechanism?

HI

c. What is the *molecularity* of step 1 in this mechanism?

Bimolecular

21. The following initial rate data were collected for the reaction 2NO + 2H2 → N2 + 2H2O.

Experiment	[NO], (M)	[H2], (M)	Initial Rate(mol/L·s)
1	0.60	0.15 doubles	0.076 doubles
2	0.60	0.30	0.15
3	doubles 0.60	0.60	quadruples 0.30
4	1.20	0.60	1.21
5	0.30	0.60	0.076

a. What order is this reaction with respect to NO?
when NO doubles, rate goes up by 4 = 2nd order

b. What order is this reaction with respect to H_2?
when H_2 doubles, rate doubles = 1st order

c. What is the value of the rate constant *k*?

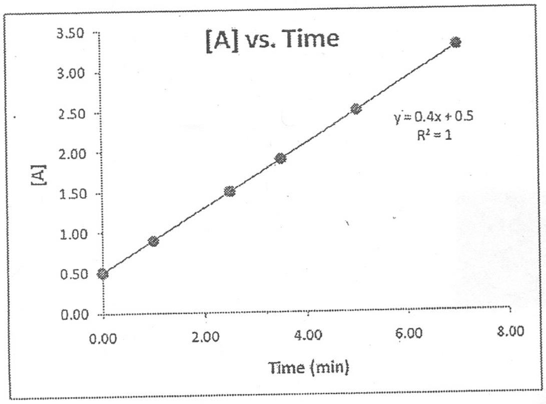
k = 1.407

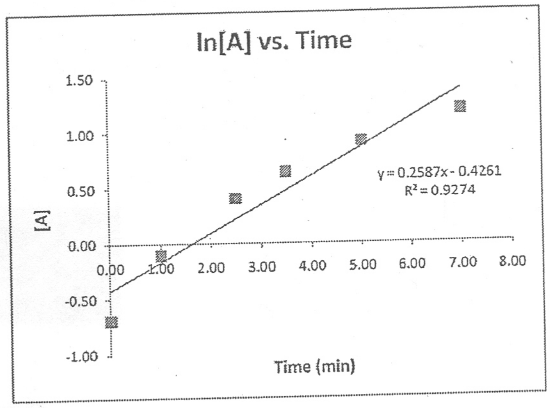
d. Write the rate law for this reaction.

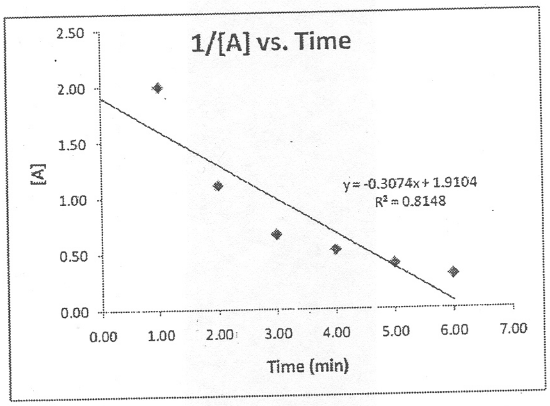
Rate = k [NO]^2[H_2]

$$k = \frac{Rate}{[NO]^2[H_2]} = \frac{0.076}{[0.6]^2[.15]} = 1.407$$

22. The following graphs represent kinetic data collected for the reaction A → B + C.







a. What is the order of this reaction with respect to A? Zero

b. What is the rate constant k for this reaction? -0.4

23. Consider the decomposition of NO_2 at 300 °C as shown below. This reaction has been experimentally determined to follow *second order* kinetics with a rate constant of 0.543 $M^{-1}s^{-1}$.

$$2\,NO_2\,(g) \rightarrow 2\,NO\,(g) + O_2\,(g)$$

$k = 0.543$

a. Derive the relationship between half-life ($t_{1/2}$) and k for a second order process. (For example, this relationship is $k = 0.693/t_{1/2}$ for a first order process.)

Integrated rate law for a second order process $= \frac{1}{[R]_t} - \frac{1}{[R]_0} = kt$

$0.5[R]_0 = \frac{1}{2}[R]_0 = [R]_t$ @ $t = t_{1/2}$

$\frac{1}{0.5[R]_0} - \frac{1}{[R]_0} = kt_{1/2}$

$\frac{1}{[R_0]}\left(\frac{1}{0.5} - 1\right) = kt_{1/2}$

$\frac{1}{[R]_0}(1) = kt_{1/2}$

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

b. What is the *half-life* for this 2nd order reaction assuming an initial concentration for NO_2 of 5.00 M?

$t_{1/2} = \frac{1}{(5M)(0.543\,M^{-1}s^{-1})} = \boxed{0.368/s}$