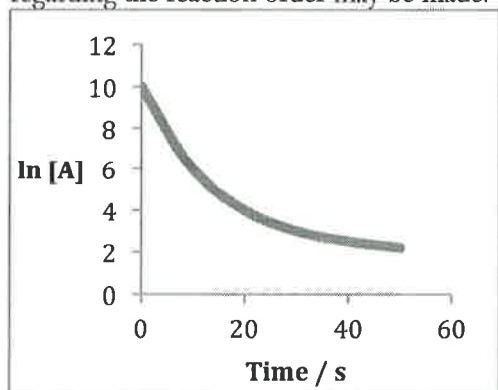# Exam 2A
Chem 1142
Spring 2013

Name: _____ KEY _____

MULTIPLE CHOICE. [5 pts ea.] Choose the best response on the scantron sheet. [**60 pts total.**]

Q1. For the gaseous reaction, $2A \longrightarrow B$ the results of a plot of ln[A] vs. time is shown below. What conclusion regarding the reaction order may be made?

a) The reaction is first order b) the reaction is not first order c) the reaction is second order
d) the reaction is not second order e) the reaction is third order

Q2. For the reaction: $A \longrightarrow B + C$, the reaction order is:
a) third order b) second order c) first order
d) zero order e) impossible to predict without more information

Q3. What are the units for k, the rate constant, in a second order reaction?
a) s^{-1} b) $M \cdot s^{-1}$ c) $M^{-1}s^{-1}$ d) s/M e) M^2

$rate = k[A]^2$

$M \cdot s^{-1} \quad M^2$

$M^{-1}s^{-1}$

Q4. What is unique about the half-life of any first order reaction?
a) the units are always s^{-1} b) the value only depend on the rate constant, k
c) the value depends only on the initial concentration of reactant
d) $\Delta[A_0]/\Delta t = 1$ e) $\Delta[A_0]/\Delta t = \frac{1}{2}$

$t_{\frac{1}{2}} = \frac{0.693}{k}$ (1st order)

Q5. The Arrhenius equation may be used to calculate the activation energy from the slope of a line plotted with what parameters?
a) ln k vs. 1/Temperature b) ln k vs. 1/time c) $1/k$ vs. Temperature
d) $1/k$ vs. 1/time e) ln k vs. e^{-T}

$k = A \cdot e^{-E_a/RT}$

$\ln k = \ln(A \cdot e^{-E_a/RT})$

$\ln k = -\frac{E_a}{R} \cdot \frac{1}{T} + \ln A$

$y = m \quad x + b$

Q6. The reaction: $A + 2B \longrightarrow C$ is thought to proceed via the mechanism:
$$B + B \longrightarrow I \quad (slow) \quad \star \text{ Rate Determining step}$$
$$I + A \longrightarrow C \quad (fast) \quad rate = k[B][B] = k[B]^2$$
The predicted rate law for this process would be:
a) rate = $k[A][B]^2$ b) rate = $k[A][2B]$ c) rate = $k[2B]$
d) rate = $k[B]^2$ e) rate = $k[I][A]$

Q7. The chemical equilibrium: $2A(s) + 3B(g) \rightleftharpoons 4C(g)$ has an equilibrium constant equal to:

a) $K_c = \dfrac{[4C]}{[2A][3B]}$

b) $K_c = \dfrac{[2A][3B]}{[4C]}$
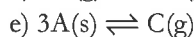
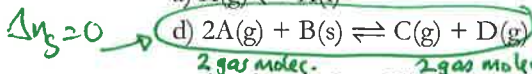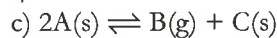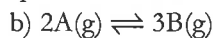c) $K_c = \dfrac{[C]^4}{[A]^2[B]^3}$ ✗

d) $K_c = \dfrac{[C]}{[A][B]}$

e) $K_c = \dfrac{[C]^4}{[B]^3}$ ✓

solids have effective conc of 1

given: $K_c = K_p (RT)^{\Delta n_g}$

if $\Delta n_g = 0$, then $K_c = K_p$

Q8. For which of the following chemical equilibria would K_c equal K_p?

a) $A(g) \rightleftharpoons A(l)$
b) $2A(g) \rightleftharpoons 3B(g)$
c) $2A(s) \rightleftharpoons B(g) + C(s)$

$\Delta n_g = 0$ → d) $2A(g) + B(s) \rightleftharpoons C(g) + D(g)$
e) $3A(s) \rightleftharpoons C(g)$

2 gas moles.   2 gas moles

Q9. For the chemical equilibrium: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, which direction would the equilibrium shift if the total reaction volume was decreased? $V\downarrow \boxed{P\uparrow}$ ⇒ will shift to reduce tot p (⇒ fewer molec gas)

a) Left
b) No-Change
c) Right
d) Not enough information given

⇒ shift RHS to convert 4 molec gas → 2 molec gas

Q10. For the chemical equilibrium: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, which direction would the equilibrium shift if some NH_3 was removed? "Wants" to make more NH_3

a) Left
b) No-Change
c) Right
d) Not enough information given

Q11. How does the addition of a catalyst affect the equilibrium constant? Speeds up fwd + rvse rxn ⇒ K_c / K_p is unchanged

a) Increases its value
b) No change to its value
c) Decreases its value
d) Impossible to answer without more information

Q12. If the equilibrium constant *increases* as the reaction temperature is increased from 10.0 °C to 20.0 °C, then:

a) temperature is acting as a catalyst ←
b) the reaction is exothermic
c) the activation energy is being lowered ← CRAZY! → d) the activation energy is being raised
e) the reaction is endothermic

cannot change E_a by changing T...

$$\text{"heat"} + R \rightleftharpoons P \qquad ; \Delta H° = +ve \ (\text{endothermic})$$

inc. T is like adding heat ⇒ will shift to RHS.

this increases K (since "heat" isn't a term in K_c expression and $K_c = \dfrac{[P]\uparrow}{[R]\downarrow}$)

Short Response.

Show ALL work to receive credit.

Q13. [20 pts.] Consider the reaction

$$X + Y \longrightarrow Z$$
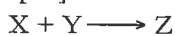
These data are obtained at 360 K:

Experiment #	[X] (M)	[Y] (M)	Initial Rate (M/s)
1	0.40	0.60	4.064
2	0.20	0.60	1.016
3	0.40	0.30	0.508

(a) Determine the rate law, and the value of the rate constant.
(b) Determine the initial rate when the concentration of X is 0.30 M and that of Y is 0.40 M.

(a) $\text{rate} = K[X]^a[Y]^b$

$$\frac{\text{rate (1)}}{\text{rate (2)}} = \frac{K[0.40M]^a[0.60M]^b}{K[0.20M]^a[0.60M]^b} = \frac{4.064 M/s}{1.016 M/s}$$

$$\Rightarrow \left(\frac{[0.40M]}{[0.20M]}\right)^a = 4.000$$

$$\Rightarrow 2.0^a = 4.000$$

$$\Rightarrow a = 2 \text{ (by inspection)}$$

$$\frac{\text{rate (1)}}{\text{rate (3)}} = \frac{K[0.40M]^a[0.60M]^b}{K[0.40M]^a[0.30M]^b} = \frac{4.064 M/s}{0.508 M/s}$$

$$\Rightarrow 2.0^b = 8.000$$

$$\Rightarrow b = 3 \text{ (by inspection)}$$

$$\left\{ or \dots \quad \log 2.0^b = \log 8.000 \right.$$
$$\left. \Rightarrow b \cdot \log 2.0 = \log 8.000 \Rightarrow b = \frac{\log 8.000}{\log 2.0} = 3.0 \right\}$$

$$\Rightarrow \text{rate} = K[X]^2[Y]^3$$

$$K = \frac{\text{rate}}{[X]^2[Y]^3}$$

substituting XPT 1 data for k $\rightarrow$

$$K = \frac{4.064\ M \cdot s^{-1}}{[0.40M]^2[0.60M]^3}$$

$$= 117.59\ M^{-4}s^{-1}$$

$$= 1.2 \times 10^2\ M^{-4}s^{-1} \text{ (2sf.)}$$

note the UNITS!!
numerator : Ms^{-1} ⎫ ratio = $M^{-4}s^{-1}$
denominator : M^5 ⎭

(b) $\text{Rate} = K[X]^2[Y]^3$

$$= 117.59\ M^{-4}s^{-1} \times [0.30M]^2 \times [0.40M]^3$$

$$= 0.677\ Ms^{-1}$$

$$= 0.68\ Ms^{-1} \text{ (2sf.)}$$

Q14. [20 pts.] K_p for the reaction: $I_2(g) \rightleftharpoons 2I(g)$ is 0.10 at a temperature of 203 °C. Imagine you started with a mixture of $I_2(g)$ and $I(g)$ with partial pressures of 1.00 atm and 1.00 atm respectively. Calculate the total pressure of the system when it reached equilibrium.

$$Q_p = P_I^2/P_{I_2} = 1.00 \quad \Rightarrow \quad Q_p > K_p, \text{ so shifts LHS.}$$

$$\underline{I_2(g) \rightleftharpoons 2I(g)}$$

	I_2	I
I	1.00 atm	1.00 atm
C	$+x$	$-2x$
E	$(1+x)$	$(1-2x)$

$$K_p = \frac{P_I^2}{P_{I_2}} \text{ @ eq}^m$$

$$\Rightarrow 0.10 = \frac{(1-2x)^2}{(1+x)}$$

$$\Rightarrow 0.10(1+x) = (1-2x)^2 = 1-4x+4x^2$$

$$\Rightarrow 0.10+0.10x = 1-4x+4x^2$$

$$\Rightarrow 4x^2 - 4.10x + 0.90 = 0$$

$$\Rightarrow x = \frac{4.10 \pm \sqrt{(-4.10)^2 - 4(4)(0.90)}}{2.4}$$

$$= \frac{4.10 \pm 1.55}{8} = 0.707 \text{ or } \boxed{0.318}$$
↑
physically impossible, since P_I would be negative!!

$$\Rightarrow P_{I_2} = 1.318 \text{ atm} \atop P_I = 0.363 \text{ atm} \Big\} \quad P_{TOT} = P_{I_2} + P_I = 1.68 \text{ atm } (2dp)$$