VERSION **1**.
Enter your version number in the correct column on
your scan sheet (see p. 2 for details).

1. Which of the following rate equations has the **highest** order overall?

   A) $v_0 = k[A]^2[B]^{-1}[C]^{-1}$ 2-1-1=0
   B) $v_0 = k[A][B]^0[C]$ 1+1+0=2
   C) $v_0 = k[A]^2[B]^{-2}[C]$ 2-2+1=1
   D) $v_0 = k[A][B][C]^0$ 1+1+0=2
   E) $v_0 = k[A]^{-1}[B]^2[C]^2$ 2+2-1=3

The overall reaction order is the sum of the individual reaction orders.

2. A phosphate buffer ($H_2PO_4^-$/HPO_4^{2-}) has a pH of 8.3. Which of the following changes will cause the **pH to increase**?

   A) adding a small amount of dilute hydrochloric acid
   B) dissolving a small amount of Na_2HPO_4
   C) dissolving a small amount of sodium chloride
   D) adding a small amount of water
   E) dissolving a small amount of NaH_2PO_4

In order to increase pH the solution needs to become more basic, a base needs to be added, namely Na_2HPO_4. NaH_2PO_4 (answer (E)) is the conjugate acid form and would make the reaction more acidic (lower pH) as would addition of HCl. NaCl is a neutral salt and would have no effect on pH and small amounts of water also does not affect the pH of a buffer.

3. You have 500. mL of a solution containing 0.20 M acetic acid (CH_3COOH) and 0.30 M sodium acetate (CH_3COONa). What will be the pH of the solution that results after the addition of 20.0 mL of 1.00 M NaOH(aq)?

Data: K_a (CH_3COOH) = 1.8×10^{-5}

First we must determine the amount of CH_3COOH and CH_3COO^- present in solution and the amount of NaOH we're adding.

A) 5.07

B) 4.84

C) 4.76

D) 4.92

E) 4.51

$n_{CH3COOH} = C_{CH3COOH}V = 0.20$ M x 0.500 L
$= 0.10$ mol
$n_{CH3COO-} = C_{CH3COO-}V = 0.30$ M x 0.500 L
$= 0.15$ mol
$n_{OH-} = C_{OH-}V_{OH-} = 1.00$ M x 0.020 L $= 0.020$ mol

The strong base will react with the weak acid of the buffer. Since the change in pH is small, and the solution remains within the buffer range, the amount of NaOH must be limiting.

$$CH_3COOH + NaOH \rightarrow H_2O + CH_3COONa$$

Before	0.10	0.020	0	0.15
After	0.08	0	0.020	0.17

$pH = pK_a + \log (n_{CH3COO-} / n_{CH3COOH})$
$= -\log(1.8 \times 10^{-5}) + \log [(0.17) / (0.08)] = 5.07$

4. Two students have 500. mL of a casein solution and access to 10.0 % (by weight) HCl(aq). They combine 1.0 mL of the 10.0% HCl with 2.5 L of distilled water, then add 8.3 mL of this solution to casein, observing the protein precipitate from solution. What is the **isoelectric point** of casein? (Assume the casein has no influence on the pH of the solution and the density of the HCl(aq) is 1.00 g mL^{-1})

First we must determine the concentration of HCl in mol L^{-1}.

A) 4.75

B) 4.63

C) 4.71

D) 4.68

E) 4.81

Assume 1.000 L of 10.0% HCl solution.
If the solution has a density of 1.00 g mL^{-1}, 1000. mL would weigh 1000. g. If it is 10.0% by weight HCl, then 100.0 g (10.0 % of the solution's mass is HCl)

n_{HCl} = mass/ molecular weight = 100.0 g / 36.46 g mol^{-1} = 2.74_{27} mol
Since we assumed 1 liter of solution the concentration is also 2.74_{27} mol L^{-1}

But 1.0 mL of this solution is diluted to 2.5 L. So the concentration of our HCl solution becomes:

$C_1V_1 = C_2V_2$
$(2.74_{27}$ mol L$^{-1})(0.0010$ L$) = C_2(2.5$ L$)$
$C_2 = 1.0_{9708} \times 10^{-3}$ mol L^{-1}

The protein precipitated when 8.3 mL of this new HCl solution was added to 500. mL of the casein solution. So the concentration of HCl in the new solution is:

$C_xV_x = C_yV_y$ note: $C_x = C_2$ and $V_y = V_{casein} + V_x$
$(1.09708 \times 10^{-3}$ mol L$^{-1})(0.0083$ L$) = C_2(0.5083$ L$)$
$C_y = 1.7_{914} \times 10^{-5}$

The pH at this concentration is $-\log(C_y) = -\log(1.7_{914} \times 10^{-5}) = 4.75$

5. Calculate the **percent ionization** in a 0.250 M solution of benzoic acid, $C_6H_5COOH(aq)$.

Data: $K_a (C_6H_5COOH) = 6.3 \times 10^{-5}$.

C_6H_5COOH is a weak acid

A) 4.1%
B) 0.90%
C) 1.6%
D) 2.2%
E) 3.4%

|	$HA + H_2O$	$\rightarrow$	$H_3O^+ +$	A^-
I	0.250		0	0
C	-x		+x	+x
E	0.250-x		+x	+x

$\dfrac{0.250}{6.3 \times 10^{-5}} = 3968 > 100$; therefore small x approximation is valid

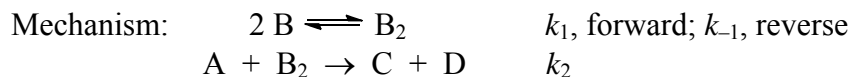$6.3 \times 10^{-5} = \dfrac{x^2}{0.250}$; $x = 3.9_{69} \times 10^{-5} = [H_3O^+]$

% dissociation = $\dfrac{x}{0.250} \times 100\% = \dfrac{3.9_{69} \times 10^{-5}}{0.250} = 1.6\%$

6. A reaction occurs in two steps with $E_{a1} = 40$ kJ mol^{-1} and $E_{a2} = 25$ kJ mol^{-1}. The first step is reversible with $E_{a-1} = 10$ kJ mol^{-1}. If the enthalpy change for the overall reaction is -27 kJ mol^{-1}, then what is the **activation energy (kJ mol^{-1})** for the reverse of step 2?

A) 65
B) 30
C) 55
D) 82
E) 47

25 kJ

10 kJ

40 kJ

40 + 27 +25 - 10 = 82 kJ

27 kJ
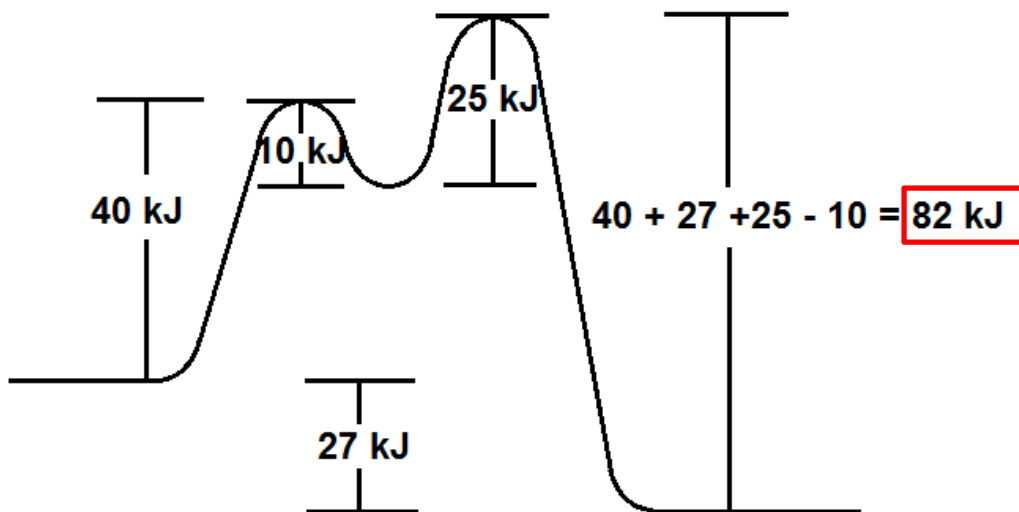
7. For the following reaction the relative magnitudes of k_1, k_{-1}, and k_2 are not known. Use the steady state approximation to derive a **rate law** for this reaction.

Overall reaction: $A + 2B \rightarrow C + D$

Mechanism: $2B \rightleftharpoons B_2$ k_1, forward; k_{-1}, reverse
$A + B_2 \rightarrow C + D$ k_2

A) $v_o = k_1[B]^2$
B) $v_o = (k_2 k_1/k_{-1})[A][B]^2$
C) $v_o = k_1[A][B]^2 / \{(k_{-1}/ k_2) + [A]\}$
D) $v_o = (k_1/k_{-1})[B]^2$
E) $v_o = k_2[B_2]$

Assumptions of steady-state approximation:

1) $v_o = k_2[A][B_2]$
2) $[B_2]$ is constant

Rate of production of $[B_2] = k_1[B]^2$
Rate of disappearance of $[B_2] = k_{-1}[B_2] + k_2[A][B_2]$

Since the $[B_2]$ is constant, Rate of production of $[B_2]$ = Rate of disappearance of $[B_2]$
$k_1[B]^2 = k_{-1}[B_2] + k_2[A][B_2]$
$k_1[B]^2 = (k_{-1} + k_2[A])[B_2]$
$[B_2] = k_1[B]^2 / (k_{-1} + k_2[A])$ … sub back into 1)

$v_o = k_2[A]k_1[B]^2 / (k_{-1} + k_2[A])$ … cancelling out k_2 we get
$v_o = k_1[A][B]^2 / \{(k_{-1}/ k_2) + [A]\}$

8. Choose the **FALSE** statement about reactions.

A) A transition state sits at higher energy than a reaction intermediate involved in the same step.
B) A transition state species has partially formed and/or partially broken bonds.
C) A reaction intermediate is produced and used up during the reaction.
D) Reaction intermediates have fully formed bonds.
E) The transition state of the rate-determining step must be the highest energy point in a given reaction progress diagram.

The rate limiting step has the highest activation energy for the given step but may not be the highest point overall. See below, the transition state of Step 1 is rate-limiting but is not the highest energy point.

4   Last Page
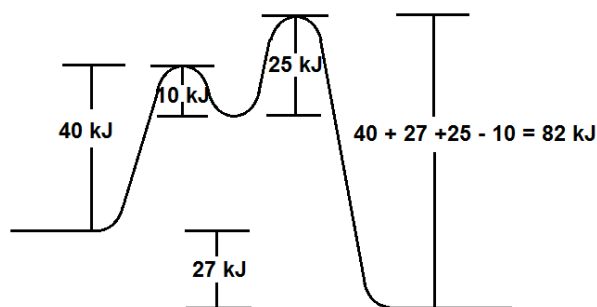
9. Which of the following graphs is **NOT linear** for the disappearance of reactant A?

A) ν vs [A] for a second order reaction
B) dν/dt vs [A] for a first order reaction
C) dν/dt vs [A] for a zero order reaction
D) ν vs [A] for a first order reaction
E) ν vs [A] for a zero order reaction

A ν vs [A] B dν/dt vs [A] C dν/dt vs [A] D ν vs [A] E ν vs [A]
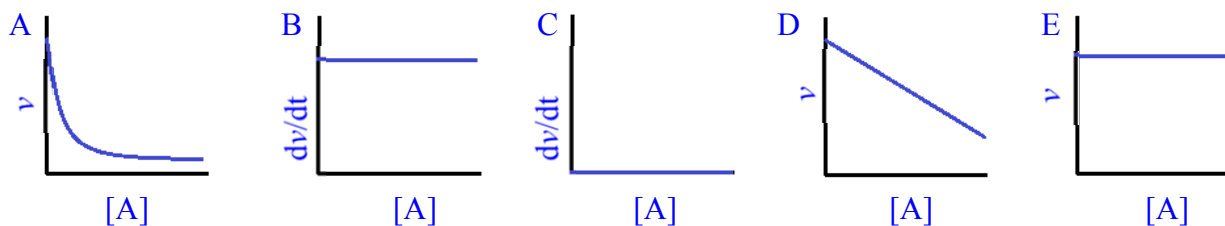
10. According to collision theory, the **increase in the rate constant** with increasing temperature is mostly because

A) the activation energy decreases with increasing temperature, allowing a greater proportion of reactants to participate in effective collisions. Activation energy remains unchanged by temperature. A catalyst lowers activation energy.
B) the fraction of the collisions having the proper orientation for reaction increases with increasing temperature. Increased temperature does not increase the *fraction* (percentage) of molecules properly aligned for reaction.
C) the pressure of the reagents increases with increasing temperature. Pressure would increase but this has nothing to do with collision theory and does not explain an increase in the rate constant.
D) the enthalpy change for most reactions is negative. Both exo and endothermic reactions have increased rate constants at increasing temperature, so the statement is false and does not account for exothermic reactions.
E) the fraction of the collisions having sufficient energy to react increases with increasing temperature.

11. Given the following data, what is the **rate law** for the reaction below? Note: D is the only substance that absorbs light at the wavelength used.

$$A + 2\,B + C \rightarrow 2\,D + E$$

Trial	[A] M	[B] M	[C] M	Time (s)	Absorbance of D
1	0.10	0.20	0.10	10	0.01
				20	0.02
2	0.10	0.40	0.10	10	0.04
				20	0.08
3	0.20	0.20	0.10	10	0.02
				20	0.04
4	0.10	0.20	0.20	10	0.01
				20	0.02

A) $v_0 = k[A][B]^2[C]$
B) $v_0 = k[A][B][C]$
C) $v_0 = k[B]^2$
D) $v_0 = k[A][B]^2$
E) $v_0 = k[B]^2[C]$

$v_0 = k[A]^m[B]^n[C]^y$ and absorbance of D is directly related to [D]
v_0 = change in absorbance in D/change in time
For reactant A, use trial 1 and 3
$v_0(3)/v_0(1) = (0.04/20)/(0.02/20) = 2$
$= k(0.20)^m(0.20)^n(0.10)^y/(0.10M)^m(0.20)^n(0.10)^y$
$= (2)^m$ → m = 1

For reactant B, use trial 1 and 2
$v_0(3)/v_0(1) = (0.08/20)/(0.02/20) = 4$
$= k(0.10)^m(0.40)^n(0.10)^y/(0.10M)^m(0.20)^n(0.10)^y$
$= (2)^n$ → n = 2

For reactant C, use trial 1 and 4
$v_0 4)/v_0(1) = (0.02/20)/(0.02/20) = 1$
$= k(0.10)^m(0.20)^n(0.20)^y/(0.10M)^m(0.20)^n(0.10)^y$
$= (2)^n$ → n = 0

12. A catalyst **increases** the rate of a reaction by

A) decreasing the temperature at which the reaction will proceed spontaneously. A catalyst does not affect the spontaneity of a reaction just how quickly it occurs (namely the rate).
B) making the reaction more exothermic. A catalyst has no effect on the enthalpy of a reaction, only the rate.
C) shifting the equilibrium position of the reaction. A catalyst has no effect on equilibrium position, it just allows it to reach equilibrium faster.
D) lowering the activation energy of the reaction.
E) increasing the activation energy for the reverse reaction. A catalyst lowers Ea equally for both the forward and reverse direction.

13. For the reaction below,

$$N_2O(g) \rightarrow N_2(g) + \tfrac{1}{2} O_2(g)$$
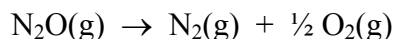
the rate constant at 1050 K is 3.4 s^{-1} and at 1384 K the reaction is 1000 times faster. If the reaction were carried out at 1050 K with a catalyst, what **activation energy (kJ mol^{-1})** would allow the reaction to run 1000 times faster at that temperature?

A) 210
B) 220
C) 170
D) 190
E) 230

$$k = Ae^{-(Ea/RT)}; \qquad \ln k_1 = -E_a/(RT_1) + \ln A$$
$$\ln k_2 = -E_a/(RT_2) + \ln A$$
$$\ln k_1/k_2 = -E_a/R \times (1/T_1 - 1/T_2)$$
$$\ln 1/1000 = -E_a /8.314 \times (1/1050 - 1/1384)$$
$$E_a = 250000J$$

$$k = Ae^{-(Ea/RT)}; \qquad \ln k_1 = -E_{a1}/(RT) + \ln A$$
$$\ln k_2 = -E_{a2}/(RT) + \ln A$$
$$\ln k_1/k_2 = 1/(RT) \times (E_{a2} - E_{a1})$$
$$\ln 1/1000 = 1 /(8.314 \times 1050) \times (E_{a2} - 250000J)$$
$$E_{a2} = 190000J = 190 \text{ kJ}$$

14. Calculate the **half-life (s)** for the reaction below at 53°C

$$(NH_3)_5CoNC\text{-}R^{3+}(aq) + OH^-(aq) \rightarrow (NH_3)_5CoNHCOR^{2+}(aq)$$
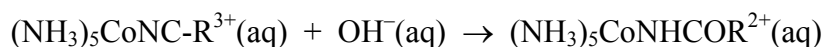
if $k = 5.8 \times 10^6$ ~~M^{-1}~~ s^{-1} at 25°C and $E_a = 92$ kJ mol^{-1}. Note there was a typo that is nonsense. If the units were in fact $M^{-1}s^{-1}$, the reaction would be second order and the half-life would be impossible to calculate since it would constantly change based on concentration.

A) 5.0×10^{-9} s
B) 7.3×10^{-6} s
C) 9.0×10^{-7} s
D) 1.3×10^{-8} s
E) 2.5×10^{-7} s

$$k = Ae^{-(Ea/RT)}; \qquad \ln k_1 = -E_a/(RT_1) + \ln A$$
$$\ln k_2 = -E_a/(RT_2) + \ln A$$
$$\ln k_1/k_2 = -E_a/R \times (1/T_1 - 1/T_2)$$

$$\ln k_2/k_1 = E_a/R \times (1/T_1 - 1/T_2)$$

$$\ln k_2/(5.8 \times 10^6 \ s^{-1}) = 92000/8.314 \times (1/298 - 1/326)$$
$$k_2 = 1.4_{08} \times 10^8 \ s^{-1}$$
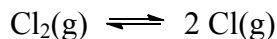
$$t_{1/2} = \ln2 / k_2$$
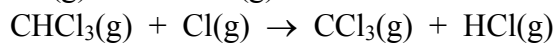$$= 4.9_2 \times 10^{-9} \ s ; \quad \text{the answer comes out as } 5.0 \times 10^{-9} \ s$$

15. The gas-phase reaction between chloroform and chlorine,

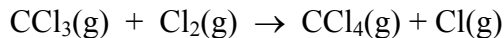$$CHCl_3(g) + Cl_2(g) \rightarrow CCl_4(g) + HCl(g)$$
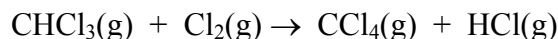
has been proposed to occur by the mechanism below. What is the **rate law** for this mechanism?

$$Cl_2(g) \rightleftharpoons 2\,Cl(g) \qquad\qquad K, \text{ fast}$$
$$CHCl_3(g) + Cl(g) \rightarrow CCl_3(g) + HCl(g) \qquad k_2, \text{ slow}$$
$$CCl_3(g) + Cl_2(g) \rightarrow CCl_4(g) + Cl(g) \qquad k_3, \text{ fast}$$

$v_o = k_2[CHCl_3][Cl]$

A) $v_o = k_2K^{0.5}[CHCl_3][Cl_2]^{0.5}$ But [Cl] is an intermediate.

B) $v_o = k_2[CHCl_3]$ $K = \dfrac{[Cl]^2}{[Cl_2]}$

C) $v_o = k_2K^{0.5}k_3[CHCl_3][Cl_2]^{0.5}[CCl_3]$

D) $v_o = k_2[CHCl_3][Cl_2]$ $[Cl]^2 = K[Cl_2]$

E) $v_o = k_3[CCl_3][Cl]$ $[Cl] = K^{0.5}[Cl_2]^{0.5}$

Sub [Cl] into initial equation.

$v_o = k_2[CHCl_3]K^{0.5}[Cl_2]^{0.5}$

$v_o = k_2K^{0.5}[CHCl_3][Cl_2]^{0.5}$

16. For the reaction

$$N_2O(g) \rightarrow N_2(g) + \tfrac{1}{2}\,O_2(g)$$
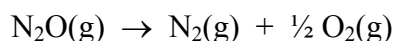
A plot of $\ln k$ versus $1/T$ gives a straight line with a slope equal to -5.1×10^3. What is the **activation energy** for the reaction?

$$k = Ae^{-E_a/RT}$$

A) 18 kJ mol^{-1}

B) 61 kJ mol^{-1}

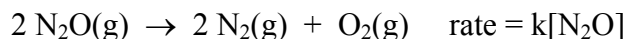C) 99 kJ mol^{-1}

D) 5.7 kJ mol^{-1}

E) 42 kJ mol^{-1}

Therefore the slope for a plot of $\ln k$ versus $1/T$ is equal to $-E_a/R$.

$-5.1 \times 10^3 = -E_a/R$

$-5.1 \times 10^3 = -E_a/(8.314)$

$E_a = 4.2_{40} \times 10^5$ J mol^{-1} = 42 kJ mol^{-1}

17. Consider the following reaction:

$$2\ N_2O(g) \rightarrow 2\ N_2(g) + O_2(g) \quad rate = k[N_2O]$$

For an initial concentration of N_2O of 0.50 M, calculate the **concentration of N_2O** remaining after 2.0 min if $k = 6.8 \times 10^{-3}\ s^{-1}$.

A) 0.44 M
B) 0.39 M
C) 0.30 M
D) 0.22 M
E) 0.15 M

Knowing the reaction is first order, the concentration can be obtained from

$$\ln([A_t]/[A_0]) = -kt \quad \text{and that 2.0 min} = 2.0 \times 60.\ s = 120sec$$

Any two data points can be taken, for example the first and last:
$$\ln([A_t]/[0.50\ M] = -(6.8 \times 10^{-3}\ s^{-1})(120\ s)$$
$$[A_t] = 0.22\ M$$

18. Choose the **TRUE** statements. A **catalyst**:

i)   provides an alternate reaction pathway
ii)   does not participate chemically in a reaction This statement is false for many catalysts like enzymes or H^+ which are catalysts that readily participate chemically.
iii)   is consumed in the reaction and therefore does not appear in the overall chemical reaction   Catalysts are not consumed in a reaction. While they may participate chemically (see previous statement) they are always regenerated by the end of the reaction mechanism.
iv)   speeds up a reaction
v)   is heterogeneous if it is in a different phase than the reactants

A) ii, v
B) ii, iv
C) i, iv, v
D) i, iii
E) ii, iii, iv

19. The following data were obtained for the disappearance of A at 25°C:

[A], M	Time, s
0.090	0
0.069	5
0.054	10
0.042	15
0.032	20
0.025	25
0.019	30

The simplest way to determine the order of the reaction is by looking at the units of the rate constant (s^{-1}) which are units for a first order reaction.

Another way is to look at the initial concentration, 0.090 M and seeing that it take approximately 14 seconds for there to be half the amount remaining ($t_{1/2}$). Then from 0.042 M at 15 s it takes another ~14 s for there to be half the concentration again ($t_{1/2}$). Constant half-lives is characteristic of a first order reaction.

What is the **order of the reaction** and what is the **rate constant (s^{-1})**?

A) First order, 0.0516
B) Second order, 0.0516
C) Second order, 0.478
D) First order, 0.478
E) Zero order, 0.0322

A third and most accurate way to determine the order is to plot or calculate the slot of [A] vs t or ln [A] vs t. If the slope of [A] vs t is linear, the reaction is 0 order. If the slope of ln [A] vs t is linear, then the reaction is 1st order.

Knowing the reaction is first order, the rate constant can be obtained from

$$\ln([A_t]/[A_0]) = -kt$$

Any two data points can be taken, for example the first and last:
$$\ln([0.019]/[0.090] = -k(30 \text{ s})$$
$$k = 0.0518 \text{ s}^{-1}$$

If all combinations were calculated, they would average 0.0516 s^{-1} with each calculation within ±0.0030 s^{-1} giving 0.0516 s^{-1} as the only reasonable answer.

20. A phase diagram for Iron (Fe) is shown below; indicate the **CORRECT** labeling of the phases/phase transitions.
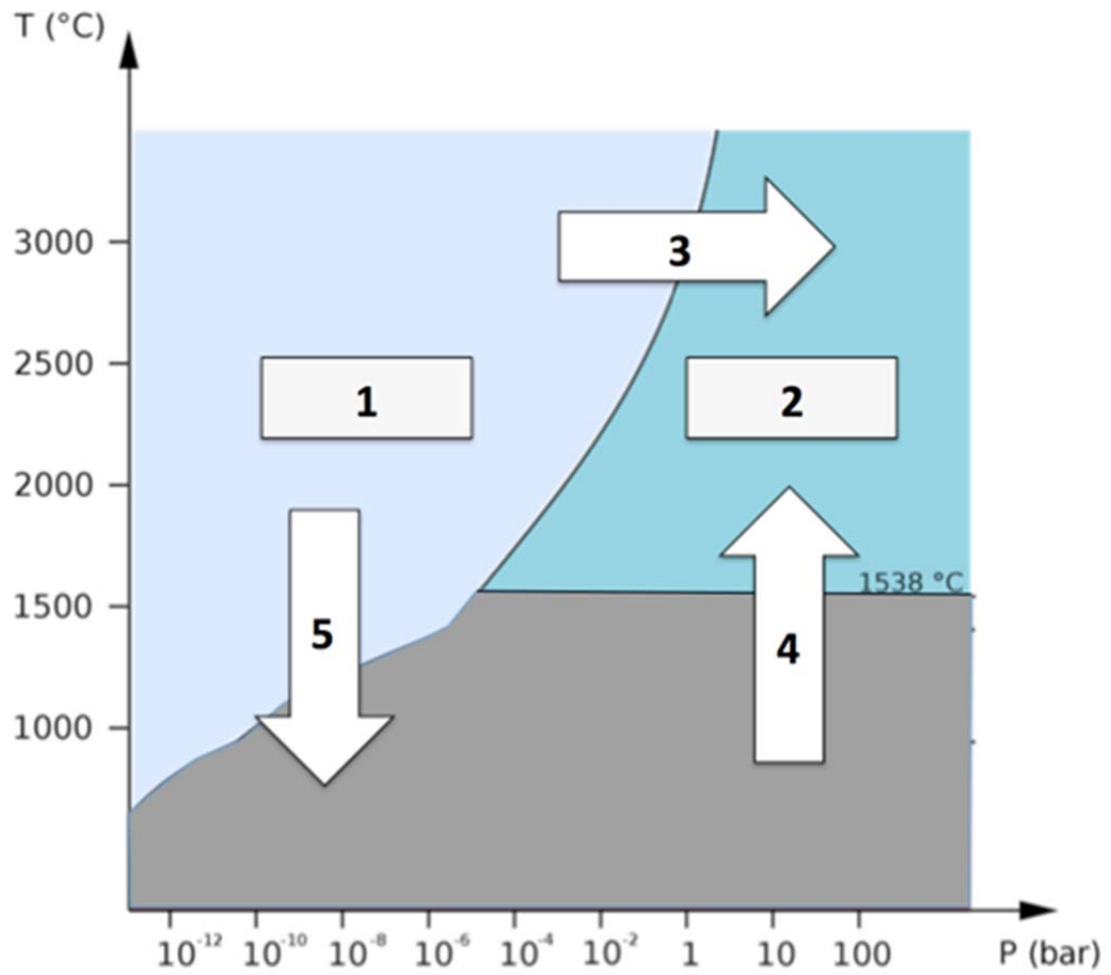
	1	2	3	4	5
A)	Vapour	Liquid	Condensation	Fusion	Deposition
B)	Vapour	Liquid	Condensation	Vaporization	Freezing
C)	Liquid	Vapour	Vaporization	Melting	Freezing
D)	Vapour	Solid	Deposition	Freezing	Melting
E)	Liquid	Vapour	Condensation	Sublimation	Melting

Note that the axes are flipped from the customary phase diagrams we've seen. This flips the positions of the solid and gaseous phases, with the phases changes also flipping accordingly.

21. Indicate the **FALSE** statement regarding the supercritical state.

A) Supercritical carbon dioxide can coexist with liquid carbon dioxide.
B) Supercritical carbon dioxide can dissolve certain organic solids.
C) Supercritical processes require significant investment of energy to compress the gas to the supercritical state.
D) Moving from liquid to vapour through the supercritical area of the phase diagram avoids crossing phase boundaries.
E) At the critical point the phase boundary between the liquid and vapour phases disappears.

Supercritical carbon dioxide is a state but not a phase. Therefore carbon dioxide can be a supercritical fluid OR a conventional liquid, but cannot be both at the same time. It is possible to have two PHASES coexist at any point along the phase boundary (i.e. solid/liquid, liquid/gas, solid/gas). However, there is no phase boundary between the liquid phase and the supercritical state.

22. Indicate the **CORRECT** graph describing the dependence of vapour pressure (mm Hg) on temperature (°C), for octane (solid line) and octanol (dotted line).
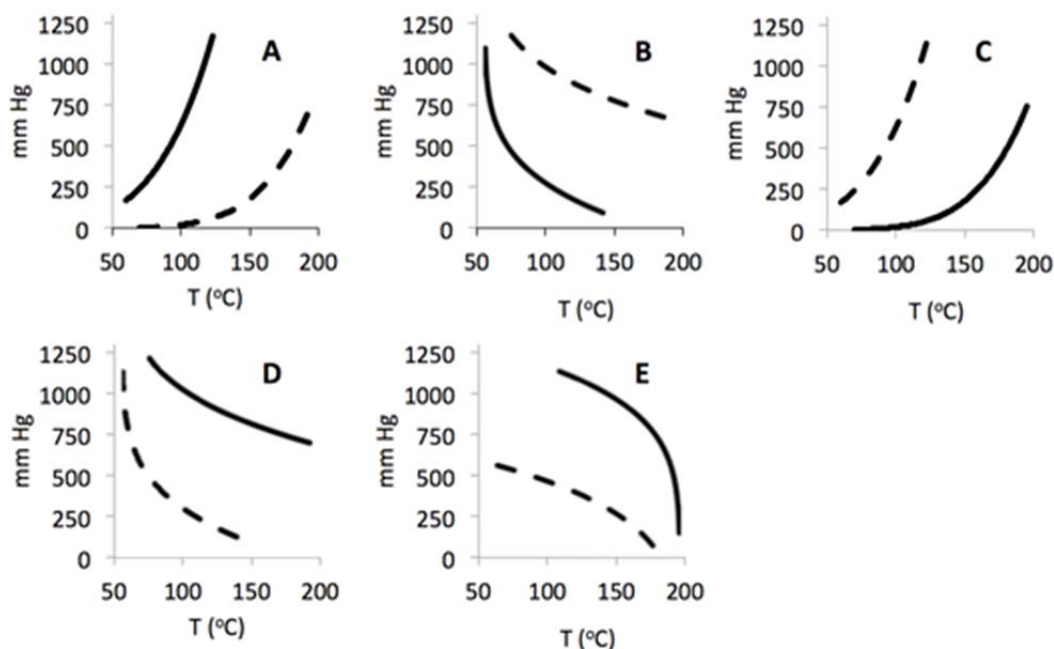
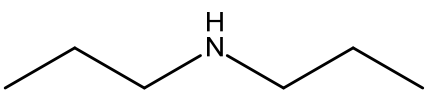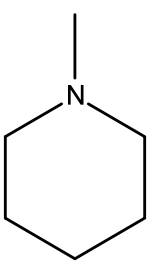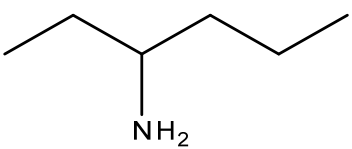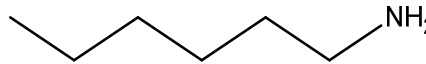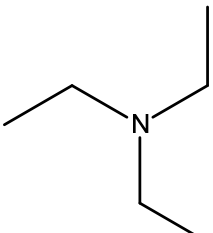A) Vapour pressure increases with temperature. The higher the temperature, the more molecules enter the gas phase and the higher the pressure. Therefore only A or C are possible. The lower the vapour pressure at a given temperature the higher the intermolecular forces (IMFs). Octanol has London Dispersion forces, dipole-dipole and hydrogen bonding IMFs while octane has only London Dispersion forces. Therefore octanol will have a lower vapour pressure than octane which is displayed in A.

23. Indicate the **FALSE** statement regarding intermolecular forces:

A) Hydrogen bonding involves a hydrogen atom plus two electronegative atoms (N, O or F).
B) Hydrogen bonds are a key factor in maintaining the three dimensional structure of most biomolecules.
C) The molecular shape of hydrocarbons can affect London (dispersion) forces, and properties such as boiling point.
D) London forces tend to decrease with increasing atomic or molecular weight, due to decreasing polarizabilities. Polarizabilities INCREASE with atomic weight.
E) London (dispersion) forces are always attractive, while interactions between permanent dipoles can be attractive or repulsive.

24. Indicate which of the following compounds has the **lowest vapour pressure** at a given temperature:

A)

B)

C)

D)

E)

The lower the vapour pressure at a given temperature the higher the intermolecular forces (IMFs). All of the molecules have approximately the same number of comparable atoms so London Dispersion forces are roughly the same. **B** and **E** have no hydrogen bonding interactions. **A** only has 1 N-H group and its in the center of the chain, making it difficult to bind to a neighbouring molecule's lone pair on N because of interference from the hydrocarbon groups (steric problems). **C** and **D** both have 2 N-H bonds but **C** has the NH$_2$ as a side-chain where it can't pack as efficiently (again steric problems). **D** has the strongest IMFs since the hydrocarbon could stack effectively with least hindrance to hydrogen bonding between NH$_2$ groups at the end of the chain.

25. Indicate the **CORRECT** attributions of intermolecular forces to acetone(l), sodium acetate(s) and acetic acid(l):

	Acetone	sodium acetate	acetic acid
i. London (dispersion)	y	y	y
ii. dipole - dipole	n	y	n
iii. hydrogen bonding	y	y	y
iv. ion - dipole	y	y	n
v. ion - ion	n	y	n

A) iii, v
B) i, v
C) i, iii, iv
D) ii, iii, iv
E) i, ii

ii.   is incorrect because acetone and acetic acid both have dipoles
iii.   is incorrect because acetic acid has hydrogen bonds
iv.   is incorrect because acetone doesn't ionize and therefore there are no ions present for an ion-dipole interaction

26. Which of the following potential energy diagrams best describes rotation around the C-C bond of **CCl₃-CBr₃**?
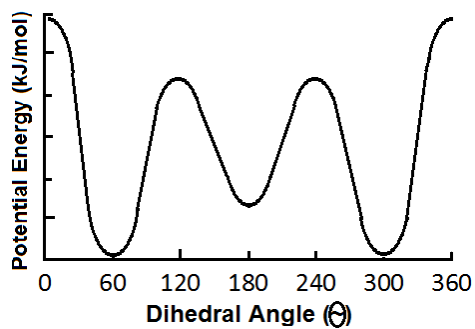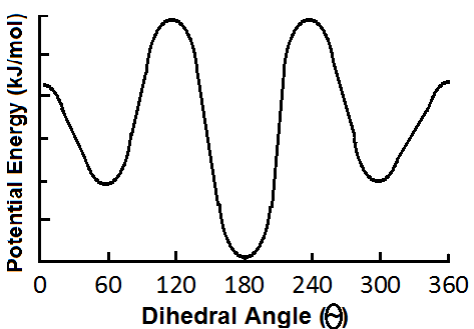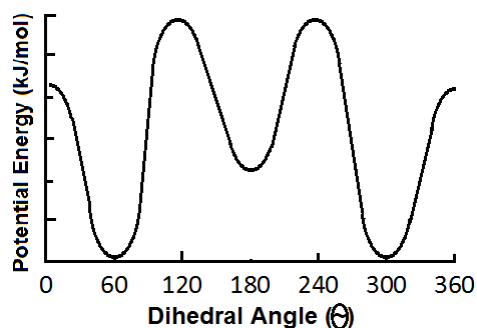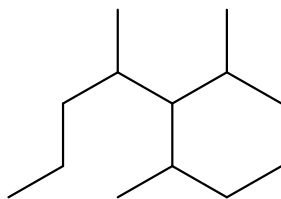
A)

B)

C)

D)

E)
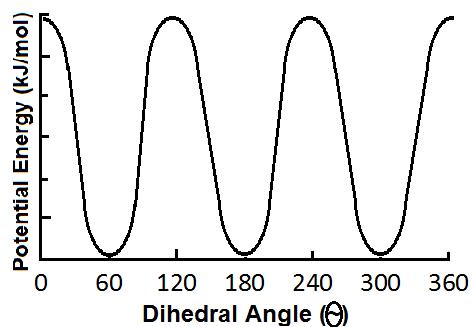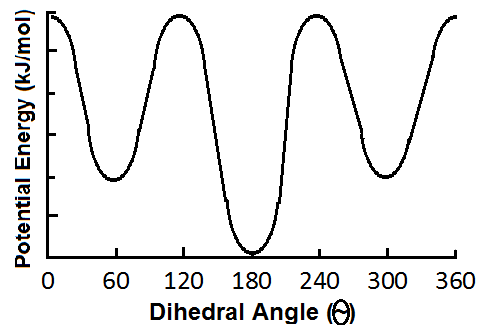
The molecules has all atoms of the same type on carbon 1 and all atoms of the same type on carbon 2 therefore there is no *gauche* or *anti* conformations and the graph would be similar to ethane which we saw in class and resembles graph A.

27. What is the **common name** for the following molecule?

A) 4,6-dimethyl-5-isopropyloctane   not named closest to first branch point
B) 4-isopropyl-3,5-dimethyloctane
C) 3,5-dimethyl-4-ethyloctane   side-chain is an isopropyl and not alphabetical
D) 2,4-dimethyl-3-*sec*-butylheptane not the longest chain
E) 5-isopropyl-4,6-dimethyloctane   not named closest to first branch point & not alphabetical

28. What is the **minimum number of carbon atoms** necessary for an alkane to possess a *sec*-butyl side chain?

A) 7
B) 11
C) 13
D) 9
E) 5

This is one example that satisfies the criteria with a total of 13 carbon atoms. It is impossible to have fewer carbon atoms because the molecule would have the *sec*-butyl group included as the parent chain either because it would be the longer chain or it would have more branch points.
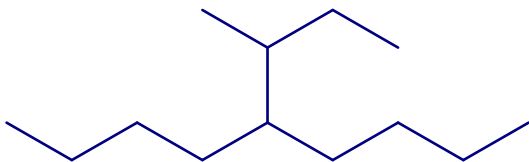
29. Identify the following **FALSE** statement.

   A) Pyridine contains a nitrogen atom.
   B) Vinyl and allyl groups contain carbon-carbon double bonds.
   C) For cyclohexane, the boat conformation is less stable than the chair conformation.
   D) The general formula for an alkane is C_nH_{2n}.   The formula is C_nH_{2n+2}.
   E) The substituents on a cyclopropane ring are eclipsed.

30. Identify the following **FALSE** statement about organic molecules and their structures.

   A) A dashed line means the bond is going backwards behind the plane.
   B) Different configurations can be achieved by rotation of single bonds.
   C) A straight line means the bond is in the plane.
   D) A solid wedged line means the bond is coming forward out of the plane.
   E) *Gauche* conformations are less stable than *anti* conformations.

Different **conformations** can be achieved by rotation of single bonds. Different configurations are obtained by altering where different atoms are bonded.