Form A B C D

Detailed Key

CHEMISTRY 102A Spring 2023
Exam III

$CO_2(g) + 2H_2O(g) \rightarrow CH_3OH(l) + 3/2\ O_2(g)$ $\Delta H_1 = -(-638\ kJ)$

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H_2 = -394\ kJ$

$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$ $\Delta H_3 = 2(-242\ kJ)$

$C(s) + 2H_2(g) + \frac{1}{2}O_2(g) \rightarrow CH_3OH(l)$

$\Delta H = \Delta H^\circ_{f,\ CH_3OH(l)} = \Delta H_1 + \Delta H_2 + \Delta H_3$

$= -240\ kJ/mol$

1 | 10
14 | 23

1. Given the following data:

$CH_3OH(l) + 3/2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$	$\Delta H = -638$ kJ
$C(\text{graphite}) + O_2(g) \rightarrow CO_2(g)$	$\Delta H = -394$ kJ
$H_2O(l) \rightarrow H_2O(g)$	$\Delta H = 44$ kJ
$H_2(g) + \frac{1}{2}\ O_2(g) \rightarrow H_2O(g)$	$\Delta H = -242$ kJ

the ΔH_f° reaction for $CH_3OH(l)$ is: $C(s) + 2H_2(g) + \frac{1}{2}O_2(g) \rightarrow CH_3OH(l)$

calculate the heat of formation, ΔH_f°, for methanol, $CH_3OH(l)$.

Use Hess's Law (work above) to solve for ΔH for this reaction.

a) −84 kJ/mol b) −125 kJ/mol c) −240. kJ/mol

d) −196 kJ/mol e) −284 kJ/mol

2 | 11
15 | 24

2. Which of the following diatomic molecules is expected to have the largest bond energy?
Hint: draw the Lewis structures.

H–H F–F O=O N≡N Cl–Cl

a) H_2 (2 e⁻) b) F_2 (14 e⁻) c) O_2 (12 e⁻) d) N_2 (10 e⁻) e) Cl_2 (14 e⁻)

N_2, with the triple bond, has the strongest bond with the largest bond energy.

3 | 12
16 | 25

3. Consider the following endothermic reaction at equilibrium:

heat + $SiHCl_3(g) + 3\ H_2O(l) \rightleftharpoons SiH(OH)_3(s) + 3\ HCl(g)$ $\Delta H = 450$ kJ

(1 mol gas → 3 mol gas)

for the mass of $SiH(OH)_3(s)$ to increase the reaction must shift right.

How many of the following four changes (I-IV) will cause the mass of $SiH(OH)_3(s)$ to increase?

Rxn shifts right to side with more moles of gas.

SR I. Increase the volume of the reaction container.

SR II. Increase the temperature. since heat is a reactant, rxn shifts right when T increases

NE III. Add $H_2O(l)$. Adding more water has no effect on the reaction.

SR IV. Remove HCl(g). Rxn shifts right to produce more HCl(g), and in turn, more $SiH(OH)_3(s)$.

a) 0 (none) b) 1 c) 2 d) 3

e) 4 [All of these changes (I-IV) will cause the mass of $SiH(OH)_3$ to increase.]

4 | 13
17 | 26

4. An equilibrium mixture for the reaction:

$2\ NO(g) + O_2(g) \rightleftharpoons 2\ NO_2(g)$ $K = \frac{[NO_2]^2}{[NO]^2[O_2]} = \frac{\left(\frac{0.12\ mol}{4.0\ L}\right)^2}{\left(\frac{0.080\ mol}{4.0\ L}\right)^2\left(\frac{0.64\ mol}{4.0\ L}\right)}$

contains 0.12 mol of NO_2, 0.080 mol of NO, and 0.64 mol of O_2 in a 4.0 L bulb at a temperature, T. What is the value of K for this reaction at this temperature?

a) 88 b) 14 c) 9.4 d) 3.5 e) 2.3

From above, $K = 14$.

$$Ca_3(PO_4)_2(s) \rightleftharpoons 3\,Ca^{2+} + 2\,PO_4^{3-}$$

	$Ca_3(PO_4)_2$	$3\,Ca^{2+}$	$2\,PO_4^{3-}$
I	—	0	0
C	$-s$	$+3s$	$+2s$
E	—	$3s$	$2s$

where s = molar solubility

From problem $[Ca^{2+}] = 4.8\times10^{-7}$ M. From ICE Table, $[Ca^{2+}] = 3s$

5 | 22
1 | 18

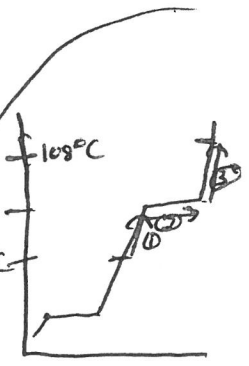
5. Excess solid $Ca_3(PO_4)_2$ is added to some water. If the equilibrium concentration of $Ca^{2+}(aq)$ is 4.8×10^{-7} *M*, what is the molar solubility of $Ca_3(PO_4)_2$ in water?

So: $3s = 4.8\times10^{-7}$ M, $s = 4.8\times10^{-7}/3$

(a) 1.6×10^{-7} mol/L b) 3.2×10^{-7} mol/L c) 4.8×10^{-7} mol/L

$s = 1.6\times10^{-7}$ mol/L

d) 1.1×10^{-19} mol/L e) 2.8×10^{-30} mol/L



There are 3 parts to this process (steeplot). $q_{TOT} = q_1 + q_2 + q_3$

6 | 23
2 | 19

6. Ethanol (C_2H_5OH) has a freezing point of −114°C and a boiling point of 78°C. The specific heat capacity of liquid ethanol is 2.5 J/g•°C and the specific heat capacity of gaseous ethanol is 1.5 J/g•°C. When a 2.00 mol sample of ethanol at 48°C is heated to 108°C, 92 kJ of heat are required for this process. Calculate the enthalpy of vaporization for ethanol (ΔH_{vap} = ?). The molar mass of ethanol is 46.07 g/mol.

$$q_{TOT} = 92{,}000\,J = \frac{2.5\,J}{g\cdot°C}\times\frac{2\,mol}{1\,mol}\times\frac{46.07\,g}{1\,mol}\times30°C + 2.0\,mol\times\Delta H_{vap} + \frac{1.5\,J}{g\cdot°C}\times2\,mol\times\frac{46.07\,g}{1\,mol}\times30°C$$

a) 81 kJ/mol b) 62 kJ/mol (c) 40. kJ/mol d) 15 kJ/mol e) 6600 J/mol

Solving: $\Delta H_{vap} = 4.047\times10^4\,J/mol = (40.\,kJ/mol)$

7 | 24
3 | 20

7. Consider the following reaction:

0 mol gas → 12 + 5 + 3 = 20 mol gas, $\Delta n = 20 - 0 = 20$ mol gas

$2\,C_7H_5N_3O_6(s) \rightarrow 12\,CO(g) + 5\,H_2(g) + 3\,N_2(g) + 2\,C(s)$ $\Delta H° = -1899$ kJ $= q_p$

Calculate $\Delta E°$ for the above reaction at 1.00 atm and 25°C.

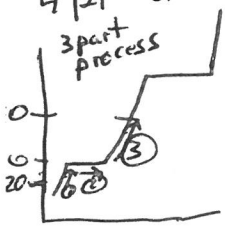
$w = -P\Delta V = -RT\Delta n = -8.3145\,J/K\cdot mol \times 298\,K \times 20\,mol = -4.955\times10^4\,J$

a) −1849 kJ b) −1899 kJ (c) −1949 kJ d) −1906 kJ e) −1892 kJ

$\Delta E = q + w = -1899\,kJ + (-49.6\,kJ) = (-1949\,kJ) = \Delta E$

9 | 25
4 | 21

8. You are given 500.0 g of ice at −20.0°C and you heat it until you have water at 40.0°C. Calculate q for the entire process of heating the H_2O from −20.0°C to 40.0°C. Pertinent data: heat capacity of water = 4.184 J/g•°C, heat capacity of ice = 2.03 J/g•°C, enthalpy of fusion of ice = 6.02 kJ/mol, molar mass of H_2O = 18.02 g/mol.



$$q_{TOT} = q_1 + q_2 + q_3 = \frac{2.03\,J}{g\cdot°C}\times500.0\,g\times20°C + \frac{500.0\,g}{18.02\,g/mol}\times6020\,J/mol + \frac{4.184\,J}{g\cdot°C}\times500.0\,g\times40°C$$

a) 10.4 kJ b) 325 kJ (c) 271 kJ d) 167 kJ e) 1.04×10^5 kJ

$q_{TOT} = 2.03\times10^4\,J + 1.67\times10^5\,J + 8.37\times10^4\,J = 2.71\times10^5\,J = (271\,kJ) = q_{TOTAL}$

9 | 26
5 | 22

9. Which of the following statements (a-d) is/are **true**?

(T) a) The internal energy of a system is equal to the sum of the potential energy and the kinetic energy of the system. → this is a definition of E.

F b) A path function is a quantity that is path ~~independent~~ dependent.

F c) When a gas expands against the surroundings, the sign associated with the work for this process is ~~positive~~ negative; system does work on surroundings in an expansion.

F d) An intensive property is a property that ~~depends~~ does not depend on the quantity of substance present.

F e) In an exothermic process, heat is ~~absorbed~~ released by the system.

Extensive properties are the ones that do depend on the amount of substance present.

Form A/B C/D

Heat added to calorimeter = $6.31\ g\ CH_4 \left|\frac{1\ mol\ CH_4}{16.04\ g}\right|\frac{802\ kJ}{mol} = 315.5\ kJ$ added

Heat capacity = $\frac{q}{\Delta T}$

Heat capacity = $\frac{315.5\ kJ\ added}{16.3^\circ C\ increase\ in\ temp} = 19.4\ kJ/^\circ C$

10 | 5
18 | 13

10. The heat capacity of a bomb calorimeter was determined by combusting 6.31 g of methane in the bomb calorimeter. The temperature changed by 16.3°C. Calculate the heat capacity of the bomb calorimeter. For methane, $\Delta E_{comb} = -802$ kJ/mol CH_4, and the molar mass of methane (CH_4) is 16.04 g/mol.

a) 311 kJ/°C b) 0.393 kJ/°C c) 3.07 kJ/°C d) 49.2 kJ/°C (e) 19.4 kJ/°C

For each salt, calculate $[Ag^+]$ where $Q = K_{sp}$. The salt will precipitate when $[Ag^+]$ is just a little larger than the calculated $[Ag^+]$; the salt requiring the smallest $[Ag^+]$ will precipitate first.

11 | 6
19 | 14

11. A solution contains 0.10 *M* SO_4^{2-}, 0.10 *M* CrO_4^{2-}, 0.10 *M* IO_3^-, and 0.10 *M* PO_4^{3-}. $AgNO_3$(aq) is added dropwise to this solution until a precipitate forms. Which of the following precipitates forms first as $AgNO_3$(aq) is added?

a) $Ag_2SO_4(s)$, $K_{sp} = 1.2 \times 10^{-5}$ $[Ag^+] = \left(\frac{1.2\times10^{-5}}{0.10}\right)^{1/2} = 1.1\times10^{-2}\ M$

b) $Ag_2CrO_4(s)$, $K_{sp} = 9.0 \times 10^{-12}$ $[Ag^+] = \left(\frac{9.0\times10^{-12}}{0.10}\right)^{1/2} = 9.5\times10^{-6}\ M$

(c) $AgIO_3(s)$, $K_{sp} = 3.1 \times 10^{-8}$ $[Ag^+] = \frac{3.1\times10^{-8}}{0.10} = 3.1\times10^{-7}\ M$ smallest

d) $Ag_3PO_4(s)$, $K_{sp} = 1.8 \times 10^{-18}$ $[Ag^+] = \left(\frac{1.8\times10^{-18}}{0.10}\right)^{1/3} = 2.6\times10^{-6}\ M$

$AgIO_3(s)$ will require the smallest $[Ag^+]$ to precipitate; it precipitates first.

12 | 7
20 | 15

12. Which of the statements (a-c) concerning the following reaction is/are **true**?

$$CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g) \qquad K = 1.0 \times 10^{-8}$$

F a) Since K<<1, at equilibrium the rate of the reverse reaction will be greater than the rate of the forward reaction. rates are equal at equilibrium

(T) b) Since K<<1, at equilibrium the reaction system will contain mostly reactants.

F c) The value of K at constant temperature depends on the amount of reactants and products that are mixed together initially. No it doesn't. K is a constant as long as T is constant.

d) Two of the above statements (a-c) are true.

e) All of the above statements (a-c) are true.

Solve the I CA ICE problem ($Q = 1.5 > K_{sp}$) to determine Ag^+ concentration at equilibrium.

	$Ag_2CO_3(s) \rightleftharpoons$	$2Ag^+$	+	CO_3^{2-}
I		1.0 M		1.5 M
C		← −1.0		−0.5
A		0		1.0
I		0		1.0
C	−s	+2s		+s
E		2s		1.0+s

13 | 8
21 | 16

13. A solution is prepared by adding 50.0 mL of 2.0 *M* $AgNO_3$ with 50.0 mL of 3.0 *M* Na_2CO_3. A precipitate of $Ag_2CO_3(s)$ forms. Calculate the equilibrium concentration of Ag^+ after precipitation is complete ($[Ag^+]_e$ = ?). K_{sp} for $Ag_2CO_3 = 8.1 \times 10^{-12}$.

Note: Volumes doubled, so initial concentrations are halved ($M_1V_1 = M_2V_2$).

(a) $2.8 \times 10^{-6}\ M$ b) $2.0 \times 10^{-12}\ M$ c) $1.4 \times 10^{-6}\ M$

$K_{sp} = (2s)^2(1.0+s) \approx 4s^2$ (assume $1.0+s = 1.0$)

d) $4.0 \times 10^{-12}\ M$ e) $9.1 \times 10^{-7}\ M$

$K_{sp} = 8.1\times10^{-12} = 4s^2$, $s = 1.4\times10^{-6}$ mol/L; assumption good.

$[Ag^+] = 2s = 2(1.4\times10^{-6}) = 2.8\times10^{-6}$ mol/L

14 | 9
22 | 17

14. A solution is prepared by adding 50.0 mL of ~~2.0 *M* $AgNO_3$ with 50.0~~ mL of 3.0 *M* Na_2CO_3. A precipitate of $Ag_2CO_3(s)$ forms. Calculate the equilibrium concentration of CO_3^{2-} after precipitation is complete ($[CO_3^{2-}]_e$ = ?). K_{sp} for $Ag_2CO_3 = 8.1 \times 10^{-12}$.

From previous problem, $s = 1.4\times10^{-6}$ mol/L

a) 1.5 *M* b) 3.0 *M* c) 0.50 *M* d) 2.0 *M* (e) 1.0 *M*

At equilibrium: $[CO_3^{2-}] = 1.0 + s = 1.0 + 1.4\times10^{-6} = 1.0\ M$

Form

A/B
C/D

CHEMISTRY 102A
Exam III

15/18
6/9

15. Consider the following four ionic compounds:

	AgCl(s)	$K_{sp} = s^2$
None are correct	$PbCl_2$(s)	$K_{sp} = 4s^3$
	Ag_3PO_4(s)	$K_{sp} = 27s^4$
	$Pb_3(PO_4)_2$(s)	$K_{sp} = 108s^5$

Next to each compound is a mathematical expression relating molar solubility to the K_{sp} value. These relationships may or may not be correct. How many of the ionic compounds has (have) the correct molar solubility to K_{sp} relationship listed for that compound? Note that s = molar solubility.

(a) 0 (None are correct.) b) 1 c) 2 d) 3 e) 4 (All are correct.)

16/19
7/10

16. A piston expands against 1.000 atm of pressure from 11.20 L to 29.10 L. In this process, 1038 J of heat are absorbed. Calculate the internal energy change for this process.

$w = -P\Delta V = -1.00\,atm(29.10 - 11.20) = -17.9\,L\cdot atm\left(\frac{101.3\,J}{L\cdot atm}\right) = -1813\,J$

$\Delta E = q + w = +1038\,J - 1813\,J = -775\,J = \Delta E$

a) −1056 J (b) −775 J c) 2851 J d) −2851 e) 1020. J

17/20
8/11

17. Consider the following reaction at 250°C:

$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g) \quad K = 0.278$

Initially 1.00 mol of N_2, 1.00 mol of H_2, and 1.00 mol of NH_3 are placed in a 2.00 L container. Which way will the reaction shift to reach equilibrium and which reagent will have the smallest concentration at equilibrium?

$Q = \frac{[NH_3]_0^2}{[N_2]_0[H_2]_0^3} = \frac{(\frac{1}{2})^2}{\frac{1}{2}(\frac{1}{2})^3} = 4.0$

Q > Ksp, so reaction shifts left to reach equilibrium.

From ICE table, $[NH_3]_e < 0.50$ M, while $[H_2]_e$ and $[N_2]_e > 0.50$ M.

a) Reaction will shift right; $[NH_3]$ will be smallest.
b) Reaction will shift left; $[H_2]$ will be smallest.
(c) Reaction will shift left; $[NH_3]$ will be smallest.
d) Reaction will shift right; $[N_2]$ will be smallest.
e) Reaction will shift right; $[H_2]$ will be smallest.

So $[NH_3]_e$ will be smallest as reaction shifts left to reach equilibrium.

18/21
9/12

18. N_2O_5(g) decomposes into O_2(g) and NO_2(g). Which of the following is a **correct** equilibrium constant expression for this decomposition reaction?

Lots of ways to balance the equation.

$N_2O_5(g) \rightleftharpoons 2NO_2(g) + \frac{1}{2}O_2(g) \quad K = \frac{[NO_2]^2[O_2]^{1/2}}{[N_2O_5]}$ ← not an answer choice, so try another balanced equation.

a) $\frac{[O_2][NO_2]^2}{[N_2O_5]}$ b) $\frac{[N_2O_5]}{[O_2][NO_2]^2}$ c) $\frac{[O_2][NO_2]}{[N_2O_5]}$

$2N_2O_5(g) \rightleftharpoons 4NO_2(g) + O_2(g) \rightarrow K = \frac{[NO_2]^4[O_2]}{[N_2O_5]^2}$ this is answer e.

d) $\frac{[N_2O_5]}{[O_2][NO_2]}$ (e) $\frac{[O_2][NO_2]^4}{[N_2O_5]^2}$

None of the other answers come from a correctly balanced equation.

Form A/B C/D

CHEMISTRY 102A Spring 2023
Exam III

19/1 23/5

19. Consider the following hypothetical reaction at 298 K:

Since the amount of C increased to reach equilibrium, the rxn shifts right.

I 1.00M 1.00M 1.00
C −x −2x +3x
E 1.00−x 1.00−2x 1.00+3x

$A(aq) + 2\,B(aq) \rightleftharpoons 3\,C(aq)$ $K = ?$

From Problem, $[C]_e = 1.96 = 1.00 + 3s$, $x = 0.32M$

where the initial concentrations are [A] = [B] = [C] = 1.00 *M*. The reaction is then allowed to proceed until equilibrium is reached, at which time the concentration of C is observed to be 1.96 *M*. What is the value of K for the above reaction (at 298 K)?

$[A]_e = 1.00 - 0.032 = 0.68M$, $[B]_e = 1.00 - 2(0.32) = 0.36m$

a) 8.0 b) 16 c) 21 (d) 85 e) 116

$$K = \frac{[C]^3}{[A][B]^2} = \frac{(1.96)^3}{0.68(0.36)^2} = 85.4 = 85$$

20/2 24/6

20. Calculate the molar solubility of solid CaF_2 ($K_{sp} = 4.0 \times 10^{-11}$) in a 0.25 *M* KF solution.

I $CaF_2(s) \rightleftharpoons Ca^{2+} + 2F^-$ $K_{sp} = 4.0 \times 10^{-11} = [Ca^{2+}][F^-]^2$
C — 0 0.25
E s 0.25+2s

a) 1.6×10^{-9} mol/L b) 2.2×10^{-4} mol/L c) $2.2 \times$ ~~10^{-6} mol/L~~

$4.0 \times 10^{-11} = s(0.25+2s)^2 \approx s(0.25)^2$, $s = 6.4 \times 10^{-10}$ mol/L Assumption good.

(d) 6.4×10^{-10} mol/L e) 0.25 mol/L

- Heat loss by water = heat gain by statue;
- heat loss by water = $\frac{4.184\,J}{g\,°C} \times 1000.0\,g \times (91.33 - 100.00) = 3.628 \times 10^4\,J$

21/3 25/7

21. A small statue, made of a pure unknown metal, has a mass of 622.5 g. The statue, initially at 26.00°C, is placed in 1.000 kg of water at 100.00°C. The final temperature of the statue and water is 91.33°C. Given that the specific heat capacity of water is 4.184 J/g•°C, what metal is the statue made of?

heat gain by statue $= 3.628 \times 10^4\,J = s \times 622.5\,g \times (91.33 - 26.00)$

a) Au, s = 0.127 J/g•°C b) Ag, s = 0.231 J/g•°C c) Zn, s = ~~0.382 J/g•°C~~

solving for s = heat capacity of statue = 0.892 J/g °C

d) Fe, s = 0.448 J/g•°C (e) Al, s = ~~0.892 J/g•°C~~

$K_p = K(RT)^{\Delta n}$, when $\Delta n = 0$, $K_p = K$ (Δn = mol gaseous products − mol gaseous reactants)

22/4 26/8

22. Consider the following four reactions:

I. $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$ (1 mol gas → 2 mol gas) $\Delta n = 2 - 1 = 1$, $K_p \neq K$

II. $2\,PBr_3(g) + 3\,Cl_2(g) \rightleftharpoons 2\,PCl_3(g) + 3\,Br_2(l)$ (5 mol gas → 2 mol gas) $\Delta n = 2 - 5 = -3$, $K_p \neq K$

III. $3\,Fe(s) + 4\,H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4\,H_2(g)$ (4 mol gas → 4 mol gas) $\Delta n = 4 - 4 = 0$, $K_p = K$

IV. $2\,H_2O(l) \rightleftharpoons 2\,H_2(g) + O_2(g)$ (0 mol gas → 3 mol gas) $\Delta n = 3 - 0 = 3$, $K_p \neq K$

For how many of these reactions does $K = K_p$?

a) 0 (none) (b) 1 c) 2 d) 3

e) 4 (All of these reactions have $K = K_p$.)

Only reaction III has $\Delta n = 0$, so $K = K_p$ for reaction III.

Form
A | B
C D

23 | 14 | 10 | 1

23. Consider the following reaction at 35°C:

$2\,NOCl(g) \rightleftharpoons 2\,NO(g) + Cl_2(g)$ $\quad K = 1.6 \times 10^{-5}$

I 2.0M 0 0
E 2.0−2x 2x x

$= \frac{(2x)^2(x)}{(2.0-2x)^2} \approx \frac{4x^3}{(2.0)^2}$

If 2.0 mol of NOCl are placed in a 1.0 L flask at 35°C, calculate the equilibrium concentration of Cl_2 ($[Cl_2]_e$ = ?).

Solving for $x = (1.6\times10^{-5})^{1/3} = 0.025\,M$ (Assumption $2-2x \approx 2$ good.)

a) $5.0 \times 10^{-2}\,M$ b) $4.0 \times 10^{-3}\,M$ c) $1.6 \times 10^{-3}\,M$

From table, $[Cl_2]_e = x = 0.025\,M$

d) $8.0 \times 10^{-3}\,M$ e) $2.5 \times 10^{-2}\,M$

24 | 15 | 11 | 2

24. The K_{sp} value for $Mg(OH)_2$ is 9.0×10^{-12}. When 100.0 mL of $4.0 \times 10^{-4}\,M$ $Mg(NO_3)_2$ is added to 100.0 mL of $2.0 \times 10^{-4}\,M$ NaOH, will a precipitate form?

$[Mg^{2+}]_0 = 2.0\times10^{-4}$, $[OH^-]_0 = 1\times10^{-4}$ (Both concentrations are halved since the volume of solution doubled.)

$Mg(OH)_2(s) \rightleftharpoons Mg^{2+} + 2OH^-$

$Q = [Mg^{2+}]_0[OH^-]_0^2 = 2.0\times10^{-4}(1.0\times10^{-4})^2 = 2\times10^{-12}$

a) Yes, because $Q = 2.0 \times 10^{-12}$ and since it is less than K_{sp}, a precipitate will form.

b) No, because $Q = 2.0 \times 10^{-12}$ and since it is less than K_{sp}, no precipitate will form.

c) Yes, because $Q = 1.6 \times 10^{-11}$ and since it is greater than K_{sp}, a precipitate will form.

d) No, because $Q = 1.6 \times 10^{-11}$ and since it is greater than K_{sp}, no precipitate will form.

Since $Q < K_{sp}$, the ion concentrations are not large enough to be at equilibrium. No precip. forms

25 | 16 | 12 | 3

25. The molar solubility for a compound is 1.0×10^{-15} mol/L. Which of the following could be the compound having this molar solubility?

a) CuS, $K_{sp} = 8.5 \times 10^{-45}$ $\quad K_{sp} = s^2$, $K_{sp} = (1\times10^{-15})^2 = 1\times10^{-30}$, no

b) AgI, $K_{sp} = 1.5 \times 10^{-16}$ $\quad K_{sp} = s^2$, $K_{sp} = 1\times10^{-30}$, no

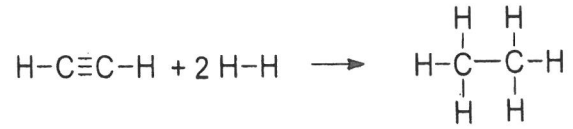
c) Ag_2S, $K_{sp} = 1.6 \times 10^{-49}$ $\quad K_{sp} = 4s^3$, $K_{sp} = 4(1\times10^{-15})^3 = 4\times10^{-45}$, no

d) Bi_2S_3, $K_{sp} = 1.1 \times 10^{-73}$ $\quad K_{sp} = 108s^5$, $K_{sp} = 108(1\times10^{-15})^5$, $K_{sp} = 1.1\times10^{-73}$, yes; Bi_2S_3 could be the compound.

e) Ag_3PO_4, $K_{sp} = 1.8 \times 10^{-18}$ $\quad K_{sp} = 27s^4 = 27(1\times10^{-15}) = 2.7\times10^{-59}$, no

26 | 7 | 13 | 4

26. Consider the following reaction and bond energies:



$\Delta H = -296$ kJ

Bond	Bond energy
C–H	413 kJ/mol
H–H	432 kJ/mol
C–C	347 kJ/mol

Bond Break
2(C–H) 2(413)
1 C≡C x
2 H–H 2(432)
1690 + x

Bonds Form
6 C–H −6(413)
1 C–C −347
−2825 kJ

Calculate the bond energy of a C≡C triple bond given the information above.

$\Delta H = \Delta H_{Break} + \Delta H_{form}$, $\Delta H = -296$ from above

a) 619 kJ/mol b) 839 kJ/mol c) 1040 kJ/mol

$-296 = 1690 + x - 2825$, x = C≡C bond energy = 839 kJ/mol

d) 562 kJ/mol e) 112 kJ/mol

A|B
C|D

A cyclic process has the same initial and final state. A state function is one whose value only depends on the difference between the initial and final states. So the change in a state function is zero for a cyclic process. Statement a is the true statement.

27|27 27|27 27|27 27. Which of the following statements is **true**?

T a) For a cyclic process, the change in a state function must be zero.

F b) At constant pressure, ~~ΔE~~ $\Delta H = q_p$.
$\Delta H = \Delta E + P\Delta V = \Delta E + RT\Delta n$, since $\Delta n \neq 0$, $\Delta H \neq \Delta E$.

F c) For $H_2O(l) \rightarrow H_2O(g)$, $\Delta H = \Delta E$ at constant P and T.

F d) The internal energy of a system ~~increases~~ decreases when the system does more work on the surroundings than heat is absorbed from the surroundings. $\Delta E = q + w$. If w is more negative than q is positive, ΔE will decrease.

F e) The energy of the university is ~~increasing~~ constant.

28|28 28|28 28. Consider the following reaction at some constant temperature:

	$2\,HF(g)$	$\rightleftharpoons$	$H_2(g)$	+	$F_2(g)$
I	0		1.0		1.0 atm
C	+2x	←	−x		−x
E	2x		1.0−x		1.0−x

$K_p = 4.0 = \frac{P_{H_2} \times P_{F_2}}{P_{HF}^2} = \frac{(1.0-x)^2}{(2x)^2}$

A container is initially filled with 1.0 atm of $H_2(g)$ and 1.0 atm of $F_2(g)$. Once equilibrium is reached, calculate the equilibrium partial pressure of $F_2(g)$.

The expression is a perfect square. Taking the square root of both sides gives:

a) 0.20 atm b) 0.40 atm c) 0.60 atm d) 0.80 atm e) 1.0 atm

$2.0 = \frac{1.0-x}{2x}$, solving: $x = 0.20$ atm; $P_{F_2} = 1.0 - 0.20 = 0.8$ atm

29|29 29|29 29|29 29. Part of the process of the gasification of coal to produce carbon dioxide and methane is shown below:

	$2\,CO(g)$	+	$2\,H_2(g)$	$\rightarrow$	$CO_2(g)$	+	$CH_4(g)$	
ΔH_f°	−110.5		0		−393.5		−74.8 kJ/mol	ΔH° = ?

$\Delta H = \sum \Delta H_{f,prods}^\circ - \sum \Delta H_{f,reacts}^\circ$

Given the following standard enthalpies of formation:

	ΔH_f° (kJ/mol)
CO(g)	−110.5
$CO_2(g)$	−393.5
$CH_4(g)$	−74.8

$\Delta H = [-393.5 + (-74.8)] - [2(-110.5)]$

$\Delta H = -468.3$ kJ $+ 221$ kJ $= -247.3$ kJ

calculate ΔH° for the above reaction.

a) −247.3 kJ/mol b) −689.3 kJ/mol c) 689.3 kJ/mol

d) 247.3 kJ/mol e) 0 kJ/mol

30. My answers for this Chemistry 102 exam should be graded with the answer sheet associated with:

a) Form A b) Form B c) Form C d) Form D e) Form E