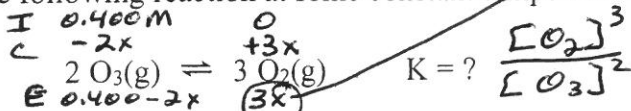


$$[O_2]_e = \frac{2.70 \text{ mol } O_2}{5.00 \text{ L}} = 0.540 \text{ M} = 3x, \quad x = 0.180 \text{ M}$$

1/14
9/22

1. Consider the following reaction at some constant temperature:



2.00 moles of $O_3(g)$ are placed into a 5.00 L rigid container and the $O_3(g)$ then reacts to reach equilibrium. At equilibrium 2.70 mol of $O_2(g)$ are present. Calculate the value of the equilibrium constant, K , for this reaction.

$$[O_3]_e = 0.400 - 2(0.180) = 0.040 \text{ M}; \quad K = \frac{(0.54)^3}{(0.040)^2} = 98$$

- a) 98 b) 490 c) 340 d) 3.4×10^4 e) 0.98

Break question 2 into 3 steps, $q_{\text{tot}} = q_1 + q_2 + q_3$

For the next two questions, consider the following data for an unknown substance X:

For Question 2, step 1 is heat $X(s)$ from -35.0 to -15.0°C . Step 2 is
 $\Delta H_{\text{vap}} = 20.00 \text{ kJ/mol}$ specific heat capacity of solid = $3.00 \text{ J/g}\cdot^\circ\text{C}$ (convert $X(s)$ at
 $\Delta H_{\text{fus}} = 5.00 \text{ kJ/mol}$ specific heat capacity of liquid = $2.50 \text{ J/g}\cdot^\circ\text{C}$ -15°C to $X(l)$
 boiling point = 75.0°C molar mass = 100.0 g/mol at -15°C , and step 3
 melting point = -15.0°C is to heat $X(l)$ from -15°C to 25°C .

2. In the heating of substance X, energy (heat) is added at a constant rate of 450.0 J/min . At this rate, how long will it take to heat 10.0 g of X from -35.0°C to 25.0°C ?

$$q_1 = 3.00 \text{ J/g}\cdot^\circ\text{C} (10.0 \text{ g}) (20.0^\circ\text{C}) = 600 \text{ J}, \quad q_2 = 0.100 \text{ mol} \times \left(\frac{5.00 \text{ kJ}}{\text{mol}} \right) = 0.500 \text{ kJ} = 500 \text{ J}$$

$$q_3 = 2.50 \text{ J/g}\cdot^\circ\text{C} (10.0 \text{ g}) (40.0^\circ\text{C}) = 1.00 \times 10^3 \text{ J}; \quad q_{\text{tot}} = 600 \text{ J} + 500 \text{ J} + 1000 \text{ J} = 2100 \text{ J}$$

$$\text{time} = \frac{2100 \text{ J}}{450.0 \text{ J/min}} = 4.67 \text{ min}$$

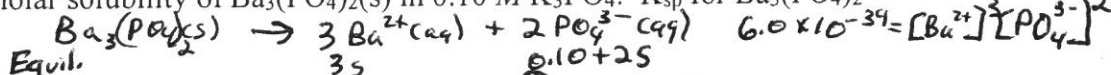
3. When the temperature of 1.000 mol of $X(g)$ is lowered from 100.0°C to form $X(l)$ at 50.0°C , 28.75 kJ of heat are released. Calculate the specific heat capacity of $X(g)$.

Step 1: $X(g, 100^\circ\text{C}) \rightarrow X(g, 75^\circ\text{C})$; Step 2: $X(g, 75^\circ\text{C}) \rightarrow X(l, 75^\circ\text{C})$; Step 3: $X(l, 75^\circ\text{C}) \rightarrow X(l, 25^\circ\text{C})$

$$28,750 \text{ J} = S_{\text{gas}}(100. \text{ g})(25^\circ\text{C}) + 1 \text{ mol}(20,000 \text{ J}) + \frac{2.50 \text{ J}}{\text{g}\cdot^\circ\text{C}} \times 100. \text{ g} \times 25^\circ\text{C}$$

$$28,750 = 2500(S_{\text{gas}}) + 20,000 + 6250, \quad S_{\text{gas}} = \frac{2500}{2500} = 1.00 \text{ J/g}\cdot^\circ\text{C}$$

4. Calculate the molar solubility of $Ba_3(PO_4)_2(s)$ in $0.10 \text{ M } K_3PO_4$. K_{sp} for $Ba_3(PO_4)_2 = 6.0 \times 10^{-39}$.



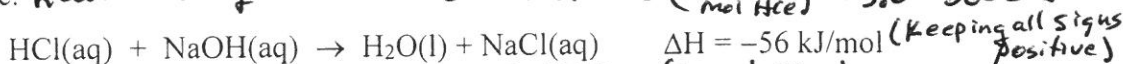
$$a) 8.9 \times 10^{-9} \text{ mol/L} \quad b) 1.1 \times 10^{-13} \text{ mol/L} \quad c) 2.8 \times 10^{-13} \text{ mol/L} \quad d) 7.7 \times 10^{-20} \text{ mol/L} \quad e) 1.1 \times 10^{-9} \text{ mol/L}$$

$$6.0 \times 10^{-39} = (3s)^3 (0.10 + 2s)^2 \approx 27s^3 (0.10)^2$$

$$\text{Solving, } s = \text{solubility} = 2.8 \times 10^{-13} \text{ mol/L}$$

Assumption good.

5. The enthalpy change for a strong acid-strong base reaction is -56 kJ/mol . Consider a coffee-cup calorimeter in which 0.10 mol of HCl is added to 0.10 mol of $NaOH$ to make 200.0 g of solution. If the initial temperature of the solution is 24.6°C and the specific heat capacity of the solution is $4.18 \text{ J/g}\cdot^\circ\text{C}$, calculate the final temperature of the reaction mixture.



$$\text{heat gain by calorimeter} = 5600 \text{ J} = 4.18 \text{ J/g}\cdot^\circ\text{C} (200. \text{ g}) \Delta T$$

- a) 25.2°C b) 26.8°C c) 28.9°C d) 29.6°C e) 31.3°C

$\Delta T = 6.7^\circ\text{C}$; since this is an exothermic reaction, the temperature of the surroundings will increase.
 $T_f = 24.6 + 6.7 = 31.3^\circ\text{C}$

2/15
10/23

3/16
11/24

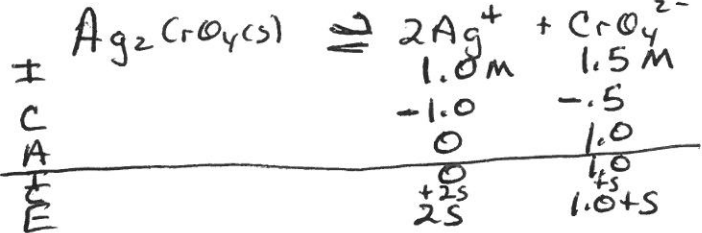
Note: I kept all quantities positive when I did this problem

4/17
12/25

5/18
13/26

Form
A/B
C/D

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$K_{sp} = 9.0 \times 10^{-12}$
Volume doubled, so initial concentrations are halved.

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Let reaction go to completion, then solve an equilibrium problem.

6/1
23/18

6. When 50.0 mL of 2.0 M $\text{AgNO}_3(\text{aq})$ is mixed with 50.0 mL of 3.0 M $\text{Na}_2\text{CrO}_4(\text{aq})$, a precipitate of $\text{Ag}_2\text{CrO}_4(s)$ forms (K_{sp} for $\text{Ag}_2\text{CrO}_4 = 9.0 \times 10^{-12}$). Once equilibrium has been established, calculate the Ag^+ concentration ($[\text{Ag}^+]_e = ?$).

$9.0 \times 10^{-12} = (2s)^2(1.0+s) \approx 4s^2, s = 1.5 \times 10^{-6} \text{ M}$ (Assumption good)

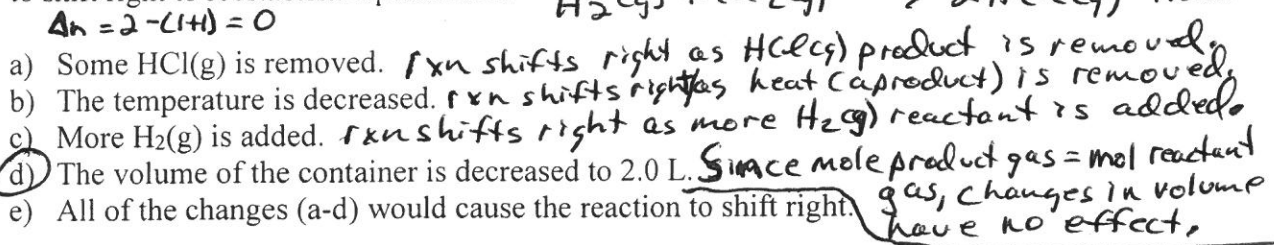
a) $1.3 \times 10^{-4} \text{ M}$ **b) $3.0 \times 10^{-6} \text{ M}$** c) $2.6 \times 10^{-4} \text{ M}$

$[\text{Ag}^+] = 2s = 2(1.5 \times 10^{-6}) = 3.0 \times 10^{-6} \text{ M}$

d) 1.0 M e) $1.5 \times 10^{-6} \text{ M}$

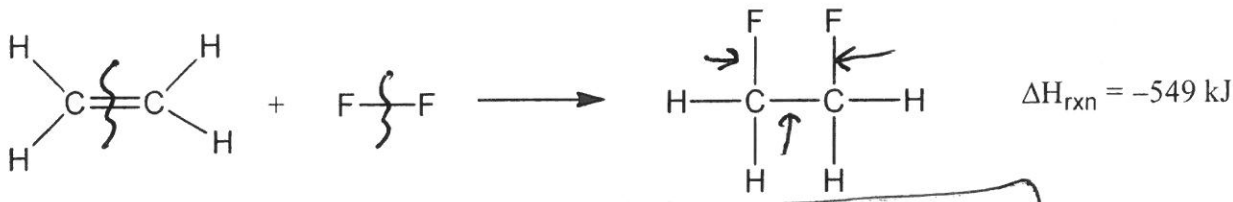
7/2
24/19

7. When $\text{H}_2(\text{g})$ and $\text{Cl}_2(\text{g})$ are reacted in a 4.0 L container, $\text{HCl}(\text{g})$ is produced in an exothermic reaction. Which of the following changes (a-d) would not cause the reaction to shift right to reestablish equilibrium?



8/3
25/20

8. Calculate the average carbon-fluorine bond energy using the following information:



Bond	Bond Energy (kJ/mol)	Break	Form
C-H	413		2 C-F
C=C	614		C-C
C-C	347		
H-H	432		
F-F	154		

$\Delta H = \sum D_{\text{broken}} - \sum D_{\text{formed}}$

$-549 \text{ kJ} = (614 + 154) - (2x + 347)$

Let $x = \text{C-F bond energy}$, solving: **$x = 485 \text{ kJ} = \text{C-F Bond Energy}$**

The average C-F bond energy is equal to:

- a) 64 kJ **b) 485 kJ** c) 970 kJ d) 128 kJ e) -970 kJ

9/4
26/21

9. Excess $\text{PbCl}_2(s)$ is added to 100.0 mL of water. Calculate the concentration of Pb^{2+} in solution ($[\text{Pb}^{2+}]_e = ?$). K_{sp} for $\text{PbCl}_2 = 1.6 \times 10^{-5}$.

$\text{PbCl}_2(s) \rightleftharpoons \text{Pb}^{2+} + 2\text{Cl}^-$ $K_{sp} = 1.6 \times 10^{-5} = [\text{Pb}^{2+}][\text{Cl}^-]^2$

$1.6 \times 10^{-5} = s(2s)^2 = 4s^3$

$s = 1.6 \times 10^{-2} \text{ M}$; from problem, $[\text{Pb}^{2+}]_e = s = 0.016 \text{ M}$

a) **0.016 M** b) 0.032 M c) 0.064 M d) 0.10 M e) $2.0 \times 10^{-3} \text{ M}$

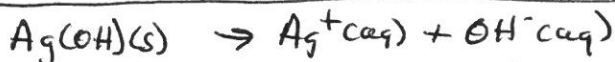
$$\Delta H = \Delta E + P\Delta V \quad (\text{at constant pressure})$$

$\Delta H = \Delta E$ when $\Delta V = 0$. This occurs when
 $\Delta n = 0$ (mol product gases = mol reactant gases),

10/19
5/14

10. ΔH° and ΔE° are approximately equal to each other in all of the following constant pressure processes except for:

- a) $\text{COCl}_2(\text{g}) \rightarrow \text{CO}(\text{g}) + \text{Cl}_2(\text{g}) \quad \Delta n = 2 - 1 = 1, \Delta H \neq \Delta E$
 b) $\text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \quad \Delta n = 0, \Delta H = \Delta E$
 c) $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightarrow 2 \text{HI}(\text{g}) \quad \Delta n = 2 - 2 = 0, \Delta H = \Delta E$
 d) $\text{CO}(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2(\text{g}) + \text{CO}_2(\text{g}) \quad \Delta n = 2 - 2 = 0, \Delta H = \Delta E$
 e) $\text{Hg}(\text{s}) \rightarrow \text{Hg}(\text{l}) \quad \Delta n = 0 - 0 = 0, \Delta H = \Delta E$



11/20
6/15

11. Consider the salt $\text{AgOH}(\text{s})$, which has $K_{\text{sp}} = 2.0 \times 10^{-8}$. How many of the following three statements (I-III) concerning the solubility of $\text{AgOH}(\text{s})$ is/are true?

- I. $\text{AgOH}(\text{s})$ is more soluble in 1.0 M HNO_3 than in pure water. $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
 II. $\text{AgOH}(\text{s})$ is more soluble in 1.0 M NH_3 than in pure water. $\text{Ag}^+ + 2\text{NH}_3 \rightarrow \text{Ag}(\text{NH}_3)_2^+$
 III. $\text{AgOH}(\text{s})$ is more soluble in 1.0 M NaOH than in pure water.
 Added H^+ removes OH^- from the K_{sp} equilibrium causing more $\text{AgOH}(\text{s})$ to dissolve. NH_3 removes Ag^+ from the equilibrium, (reaction shifts right) causing more solid to dissolve. Added OH^- will shift K_{sp} equilibrium to the left to form more $\text{AgOH}(\text{s})$.
- a) 0 (none) b) 1 c) 2
- d) 3 [All of the above statements (I-III) are true].

12/21
7/16

12. Which of the following statements is false?

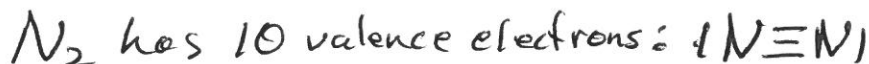
- T a) In a bomb calorimeter, ΔE can be determined.
 F b) When an endothermic reaction is performed in a calorimeter, the temperature of the calorimeter increases. Temp. decreases for an endothermic reaction.
 T c) An element in its standard state has a standard enthalpy of formation equal to zero ($\Delta H_f^\circ = 0$).
 T d) Hess's law is based on the fact that enthalpy is a state function.
 T e) In a coffee cup calorimeter, the heat released or gained at constant pressure (q_p) can be determined.

When the reactant bond strengths are overall stronger than product bond strengths, the reaction is endothermic.

13. Consider the following endothermic reaction: a
 $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{NO}(\text{g})$
 Since endothermic, the reaction has a positive ΔH value. This generally occurs when the reactant bonds are stronger than the product bonds.

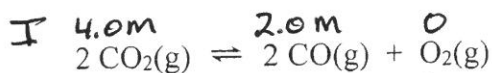
Which of the following statements is false concerning this reaction? N_2 has a triple bond, which will have a large bond energy.

- T a) N_2 should have a relatively large bond energy.
 T b) The enthalpy change for this reaction is positive.
 F c) The reactant bonds are overall weaker than the product bonds. This would lead to an exothermic rxn.
 T d) Bond energies should give a good estimate for the enthalpy change since all substances are gases.



14/5
19/10

14. Consider the following reaction at some constant temperature:



$$K = 2.0 \times 10^{-6} = \frac{[\text{CO}]^2 [\text{O}_2]}{[\text{CO}_2]^2}$$

E $4.0-2x \quad 2.0+2x \quad x$
In an experiment, 4.0 mol of CO_2 and 2.0 mol of CO are each placed into a 1.0 L rigid container. Calculate the concentration of O_2 ($[\text{O}_2]_e = ?$) once the reaction has reached equilibrium.

$$2.0 \times 10^{-6} = \frac{(2.0+2x)^2 x}{(4.0-2x)^2} \approx \frac{(2.0)^2 x}{(4.0)^2}, \quad x = 8.0 \times 10^{-6} M = [\text{O}_2]$$

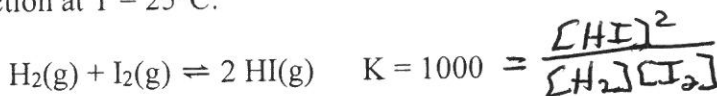
- a) $5.0 \times 10^{-7} M$ b) $1.0 \times 10^{-6} M$

- d) $4.0 \times 10^{-6} M$ e) $8.0 \times 10^{-6} M$

c) $2.0 \times 10^{-6} M$
Assumptions good.

15/6
20/11

15. Consider the following reaction at $T = 25^\circ\text{C}$:



Which of the following statements is false?

F a) The value of the equilibrium constant depends on the amounts of reactants and products that are mixed together initially. K is a constant (at constant temp).

T b) For this reaction, $K_p = K$. Since $\Delta n = 0$, $K = K_p$.

T c) At equilibrium, the rate of the forward reaction equals the rate of the reverse reaction. Always true

T d) At equilibrium, $[\text{HI}]^2 > [\text{H}_2][\text{I}_2]$. Since $K > 1$, this is true.

T e) At equilibrium, the concentration of reactants and products are constant. Always true at equilibrium

16/7
21/12

16. How many of the following (I-V) is/are intensive properties?

- I. Mass
- II. Temperature
- III. Volume
- IV. Concentration
- V. Enthalpy

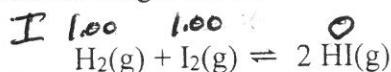
Intensive properties are ones that do not depend on the quantity of substance present. Temperature and concentration are intensive properties. Consider two air samples: 1.0 L of air molecules would have the same temperature and concentration of gases present as 1000 L of the same air. This is not the case for mass, volume, and ΔH . All of these quantities depend on the amount of substance present.

- a) 1 b) 2

- c) 3 d) 4 e) 5 (All are intensive properties.)

17/8
22/13

17. Consider the following reaction at some constant temperature:



$$K_p = 16 = \frac{P_{\text{HI}}^2}{P_{\text{H}_2} \times P_{\text{I}_2}} = \frac{(2x)^2}{(1.00-x)^2}$$

E $1.00-x \quad 1.00-x \quad 2x$
In an experiment, $\text{H}_2(g)$ at 1.00 atm and $\text{I}_2(g)$ at 1.00 atm are initially reacted. At equilibrium, what is the partial pressure of $\text{HI}(g)$?

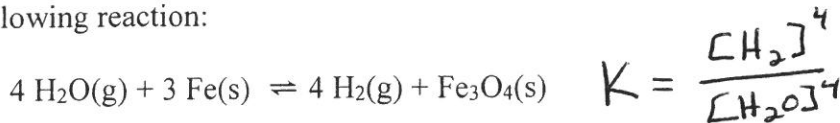
Taking the square root of both sides of the equation gives:

- a) 1.5 atm b) 0.67 atm c) 1.3 atm d) 0.75 atm e) 1.00 atm

$$4.0 = \frac{2x}{1.00-x}, \quad 2x = 4.0 - 4x, \quad 6x = 4.0, \quad x = \frac{4}{6} = \frac{2}{3} \text{ atm}$$

$$P_{\text{HI}} = 2x = 2\left(\frac{2}{3} \text{ atm}\right) = \frac{4}{3} = 1.3 \text{ atm}$$

18/23 18. Consider the following reaction:



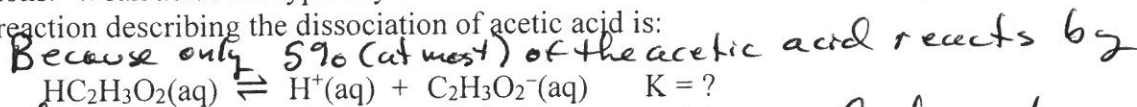
What is the correct equilibrium constant expression for this reaction?

Solids and liquids are not included in K expressions.

a) $K = \frac{[\text{H}_2\text{O}]^4 [\text{Fe}]^3}{[\text{H}_2]^4 [\text{Fe}_3\text{O}_4]}$ b) $K = \frac{[\text{H}_2][\text{Fe}_3\text{O}_4]}{[\text{H}_2\text{O}][\text{Fe}]}$ c) $K = \frac{[\text{H}_2]}{[\text{H}_2\text{O}]}$

d) $K = \frac{[\text{H}_2]^4 [\text{Fe}_3\text{O}_4]}{[\text{H}_2\text{O}]^4 [\text{Fe}]^3}$ e) $K = \frac{[\text{H}_2]^4}{[\text{H}_2\text{O}]^4}$

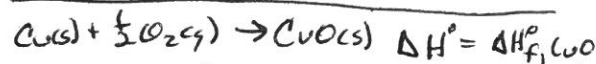
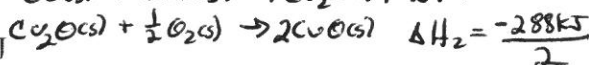
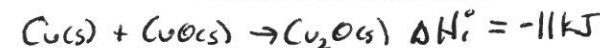
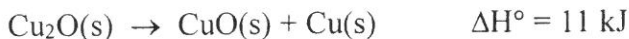
19/24 19. Earlier in the semester, we labeled weak acids as weak electrolytes. Weak electrolytes are substances where only a small quantity of the compound breaks up into ions. Weak acids are typically less than 5% dissociated in water. The equilibrium reaction describing the dissociation of acetic acid is:



this reaction, we will have mostly products at equilibrium. Given that acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) is a weak electrolyte (weak acid), which of the following is most likely the K value for the above equilibrium reaction?

one answer has $K < 1$, so it must be correct.
 a) 7.2×10^{27} b) 6.1×10^{14} c) 5.4×10^{12} d) 1.5 e) 1.8×10^{-5}

20/25 20. Given:



Calculate the standard enthalpy of formation (ΔH_f°) for $\text{CuO}(\text{s})$.

- a) 155 kJ/mol b) -155 kJ/mol c) 299 kJ/mol
 d) -299 kJ/mol e) -166 kJ/mol

This is the standard enthalpy of formation reaction for $\text{CuO}(\text{s})$ (making 1 mol of $\text{CuO}(\text{s})$ from the elements in their standard state).
 $\Delta H^\circ = -155 \text{ kJ}$

21/26 21. How many of the following five silver salts are more soluble in acidic (H^+) solution than in pure water?

- Ag₂CO₃ AgCl AgI AgBr AgClO₄

- a) 1 b) 2 c) 3 d) 4

e) 5 (All are more soluble in acidic solution than in pure water.) Cl^- , I^- , Br^- , and ClO_4^- are not more soluble in acidic solution.

FTF: hydroxide salts are also more soluble in acidic solution since the added H^+ will react with OH^- to form $\text{H}_2\text{O}(\text{l})$.

Form
A/B
C/D

Compounds with weak bonds have high potential energy, while compounds with strong bonds have low potential energy. The more stable the compound, the lower the potential energy.

22/9
14/1

22. Which of the following statements (a-c) correctly relates chemical potential energy and bond strength?

- a) Compounds with strong bonds have high potential energy.
- b) Compounds with weak bonds have low potential energy.
- c) Compounds with weak bonds have high potential energy.
- d) Statements a and b are both correct.

23/10
15/2

23. Consider the following reaction and equilibrium constant value at some temperature:



An equilibrium mixture contains 6.0 mol of $\text{N}_2\text{O}_4(\text{g})$ in a 3.0 L container. How many moles of $\text{NO}_2(\text{g})$ are present in this equilibrium mixture?

$[\text{N}_2\text{O}_4]_e = 6.0 \text{ mol} / 3.0 \text{ L} = 2.0 \text{ M}$ $8.0 = \frac{[\text{NO}_2]^2}{2.0}$ $[\text{NO}_2] = 4.0 \text{ M}$

- a) 2.0 mol
- b) 4.0 mol
- c) 6.0 mol
- d) 8.0 mol
- e) 12 mol

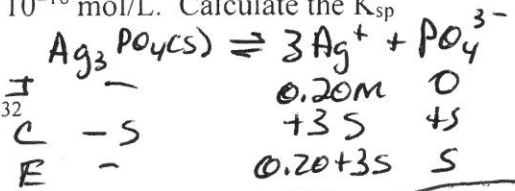
$\text{mol NO}_2 = 3.0 \text{ L} \left(\frac{4.0 \text{ mol NO}_2}{\text{L}} \right) = 12 \text{ mol NO}_2$

24/11
16/3

24. The solubility of $\text{Ag}_3\text{PO}_4(\text{s})$ in 0.20 M AgNO_3 is 2.2×10^{-16} mol/L. Calculate the K_{sp} value for Ag_3PO_4 .

$S = 2.2 \times 10^{-16} \text{ mol/L}$ (from problem)

- a) 1.8×10^{-18}
- b) 6.3×10^{-62}
- c) 4.8×10^{-32}



d) 4.4×10^{-17} e) 9.7×10^{-33}

$$K_{sp} = [\text{Ag}^+]^3 [\text{PO}_4^{3-}] = [0.20 + 3(2.2 \times 10^{-16})]^3 (2.2 \times 10^{-16}) = 1.8 \times 10^{-18} = K_{sp}$$

25/12
17/4

25. Which of the following statements is false?

- T a) The energy of the universe is constant.
- T b) Heat is a path function.
- T c) The internal energy of a system is the sum of its potential and kinetic energies.
- F d) When a gas expands against a constant external pressure, the surroundings are doing work on the system. In an expansion, the system is doing work on the surroundings.

26/13
18/5

26. Which of the following statements is false?

- $\Delta E = q + w$
- T a) The internal energy of a system decreases when more work is done by the system, than heat is flowing into the system. ΔE is negative because w is more negative than q is positive.
 - T b) At constant volume, the internal energy change for a system is equal to the amount of heat flow ($\Delta E = q_v$). $\Delta E = q + w$; $w = 0$ at constant V , so $\Delta E = q$
 - T c) The internal energy of a system increases when work is done on the system and heat is flowing into the system. $\Delta E = q + w$; ΔE is positive when both q and w are positive values.
 - F d) For the process of boiling water, the sign of q is negative.

heat + $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$
To boil water heat must be added (this is an endothermic process). q is positive for endothermic processes.

Heat added to calorimeter = $1.22 \text{ g C}_6\text{H}_5\text{COOH} \left(\frac{1 \text{ mol C}_6\text{H}_5\text{COOH}}{122 \text{ g}} \right) \left(\frac{3220 \text{ kJ}}{\text{mol C}_6\text{H}_5\text{COOH}} \right)$

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27. When 32.2 kJ added to the calorimeter, the temperature of the combustion of benzoic acid ($\text{C}_6\text{H}_5\text{COOH}$, molar mass = 122 g/mol) is commonly used to calibrate bomb calorimeters. In a bomb calorimeter, 1.22 g of benzoic acid is combusted, resulting in a temperature increase from 23.25°C to 26.47°C. Given that the energy of combustion of benzoic acid is -3220 kJ/mol, calculate the heat capacity of the calorimeter.

- increased by 3.22°C ($26.47 - 23.25 = 3.22$). A change of 3.22°C = 3.22 K.
- a) 1000 kJ/K **b) 10.0 kJ/K** c) 1220 kJ/K
d) 0.379 kJ/K e) 8.20 kJ/K
- heat capacity = $\frac{32.2 \text{ kJ}}{3.22 \text{ K}} = 10.0 \text{ kJ/K}$

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28. Consider the following solutions prepared in two separate beakers:
 $\text{PbI}_2(\text{s}) \rightleftharpoons \text{Pb}^{2+} + 2\text{I}^-$ $K_{sp} = 1 \times 10^{-8}$; $\text{PbSO}_4(\text{s}) \rightleftharpoons \text{Pb}^{2+} + \text{SO}_4^{2-}$ $K_{sp} = 2.0 \times 10^{-8}$

In each beaker, the initial concentrations are halved since the volume doubled for each beaker, calculate Q.

Beaker 1: 100.0 mL of $3.0 \times 10^{-4} \text{ M Pb(NO}_3)_2(\text{aq})$ mixed with 100.0 mL of $2.0 \times 10^{-4} \text{ M KI(aq)}$; K_{sp} for $\text{PbI}_2 = 1 \times 10^{-8}$
 $Q = [\text{Pb}^{2+}]_0 [\text{I}^-]_0^2 = 1.5 \times 10^{-4} (1.0 \times 10^{-4})^2 = 1.5 \times 10^{-12}$ $Q < K_{sp}$, so no precipitate.

Beaker 2: 100.0 mL of $3.0 \times 10^{-4} \text{ M Pb(NO}_3)_2(\text{aq})$ mixed with 100.0 mL of $2.0 \times 10^{-4} \text{ M K}_2\text{SO}_4(\text{aq})$; K_{sp} for $\text{PbSO}_4 = 2.0 \times 10^{-8}$
 $Q = [\text{Pb}^{2+}]_0 [\text{SO}_4^{2-}]_0 = 1.5 \times 10^{-4} (1.0 \times 10^{-4}) = 1.5 \times 10^{-8}$ $Q < K_{sp}$, so no precipitate.

In which of the beakers will a precipitate form?

- a) $\text{PbI}_2(\text{s})$ will form in beaker 1 and $\text{PbSO}_4(\text{s})$ will form in beaker 2.
b) $\text{PbI}_2(\text{s})$ will form in beaker 1 and no precipitate will form in beaker 2.
c) No precipitate will form in beaker 1 and $\text{PbSO}_4(\text{s})$ will form in beaker 2.
d) No precipitate will form in either beaker.

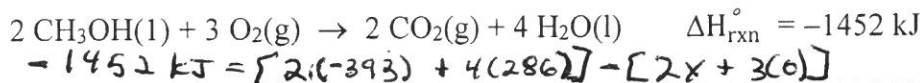
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29. Which of the following salts has the largest molar solubility in pure water?
For FeCO_3 , $K_{sp} = s^2$, $s = 4.6 \times 10^{-6} \text{ mol/L}$; For Ag_2CrO_4 , $K_{sp} = 4s^3$, $s = 1.3 \times 10^{-4} \text{ mol/L}$

- a) FeCO_3 , $K_{sp} = 2.1 \times 10^{-11}$ b) Ag_2CrO_4 , $K_{sp} = 9.0 \times 10^{-12}$
c) CaF_2 , $K_{sp} = 4s^3$, $s = 2.2 \times 10^{-4} \text{ mol/L}$ d) AgCl , $K_{sp} = s^2$, $s = 1.3 \times 10^{-5} \text{ mol/L}$
e) CaF_2 , $K_{sp} = 4.0 \times 10^{-11}$ f) AgCl , $K_{sp} = 1.6 \times 10^{-10}$

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30. Consider the reaction: $\Delta H^\circ = \sum \Delta H_{f, \text{prod}}^\circ - \sum \Delta H_{f, \text{react}}^\circ$; Let $x = \Delta H_f^\circ$ for CH_3OH



for which ΔH_f° for $\text{H}_2\text{O}(\text{l}) = -286 \text{ kJ/mol}$ and ΔH_f° for $\text{CO}_2(\text{g}) = -393 \text{ kJ/mol}$.

Calculate the standard enthalpy of formation for methanol.

Solving: $x = -239 \text{ kJ/mol} = \Delta H_f^\circ \text{ CH}_3\text{OH}$

- a) 239 kJ/mol b) -478 kJ/mol c) 478 kJ/mol **d) -239 kJ/mol** e) 0 kJ/mol

31. 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