

Work for #8 below

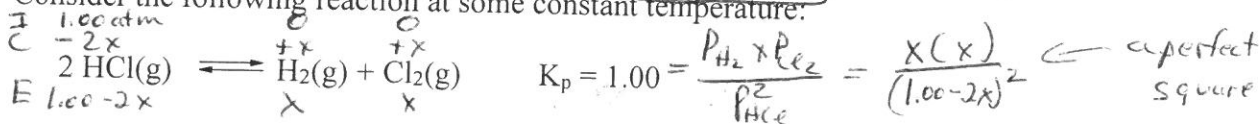
Heat for 50°C → 150°C

Step 1: $H_2O(l, 50^\circ C) \rightarrow H_2O(l, 100^\circ C)$

Step 2: $H_2O(l, 100^\circ C) \rightarrow H_2O(g, 100^\circ C)$

Step 3: $H_2O(g, 100^\circ C) \rightarrow H_2O(g, 150^\circ C)$

1. Consider the following reaction at some constant temperature:



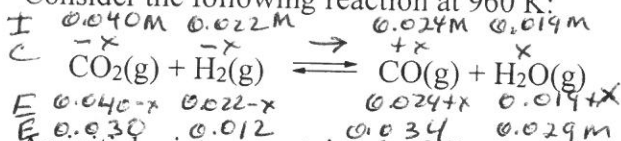
If 1.00 atm of HCl is initially reacted in a rigid container, calculate the equilibrium partial pressure of HCl.

$$1.00 = \frac{x^2}{(1.00-2x)^2} \quad 1 = \frac{x}{1.00-2x} \quad 1.00-2x=x \quad 1.00=3x \quad x=0.333 \text{ atm}$$

- a) 0.67 atm b) 1.33 atm c) 0.75 atm d) 0.50 atm e) 0.33 atm

$$P_{HCl} = 1.00 - 2(0.333) = 0.33 \text{ atm}$$

2. Consider the following reaction at 960 K:



From problem $[CO_2] = 0.030 = 0.040 - x$
 $x = 0.010 \text{ M}$

An initial mixture contains the following concentrations of reactants and products: $[CO_2] = 0.040 \text{ M}$, $[H_2] = 0.022 \text{ M}$, $[CO] = 0.024 \text{ M}$, and $[H_2O] = 0.019 \text{ M}$. Once equilibrium is established, the equilibrium concentration of CO_2 is 0.030 M. Calculate the value of K for this reaction?

$$K = \frac{[CO][H_2O]}{[CO_2][H_2]} = \frac{(0.034)(0.029)}{(0.030)(0.012)} = 2.7$$

- a) 0.52 b) 1.5 c) 0.69 d) 0.37 e) 2.7

See above for 3 steps. $q_1 = 4.18 J/g \cdot ^\circ C \times \text{mass} \times 50^\circ C$; $q_2 = 2260 J/g \times \text{mass}$; $q_3 = 2.02 J/g \cdot ^\circ C \times \text{mass} \times 50^\circ C$

3. When 385 kJ of energy is added to some water of unknown mass, the temperature increases from 50.0°C to 150.0°C. Calculate the mass of the water sample. For H_2O : boiling point = 100.0°C, specific heat capacity of $H_2O(l) = 4.18 \text{ J/g} \cdot ^\circ C$, specific heat capacity of $H_2O(g) = 2.02 \text{ J/g} \cdot ^\circ C$, $\Delta H_{vap} = 2.26 \text{ kJ/g}$.

$$q_{tot} = q_1 + q_2 + q_3 \quad 385,000 \text{ J} = 2090(\text{mass}) + 2260(\text{mass}) + 101(\text{mass})$$

- a) 31.5 g b) 150. g c) 320. g d) 540. g e) 1370 g

Solving: $\text{mass} = 150. \text{ g}$

4. Consider a solution which contains $1 \times 10^{-4} \text{ M NaF}$, $1 \times 10^{-4} \text{ M Na}_2\text{S}$, and $1 \times 10^{-4} \text{ M Na}_3\text{PO}_4$. $Pb(NO_3)_2(aq)$ is gradually added to this solution. What is the order of precipitation as $Pb(NO_3)_2$ is added?

K_{sp} for $PbF_2 = 4 \times 10^{-8}$, K_{sp} for $PbS = 7 \times 10^{-29}$, and K_{sp} for $Pb_3(PO_4)_2 = 1 \times 10^{-54}$

$PbF_2: [Pb^{2+}] = \frac{4 \times 10^{-8}}{[F^-]^2} = \frac{4 \times 10^{-8}}{(1 \times 10^{-4})^2} = 4.0 \times 10^{-4} \text{ M}$

$PbS: [Pb^{2+}] = \frac{7 \times 10^{-29}}{[S^{2-}]} = \frac{7 \times 10^{-29}}{1 \times 10^{-4}} = 7 \times 10^{-25} \text{ M}$

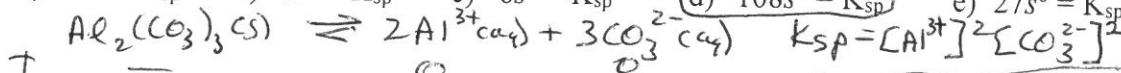
$Pb_3(PO_4)_2: [Pb^{2+}]^3 = \frac{1 \times 10^{-54}}{[PO_4^{3-}]^2} = \frac{1 \times 10^{-54}}{(1 \times 10^{-4})^2} = 1 \times 10^{-46}$
 $[Pb^{2+}] = 1 \times 10^{-16} \text{ M}$

- a) PbS precipitates first, then PbF_2 precipitates, and $Pb_3(PO_4)_2$ precipitates last.
b) PbS precipitates first, then $Pb_3(PO_4)_2$ precipitates, and PbF_2 precipitates last.
c) PbF_2 precipitates first, then $Pb_3(PO_4)_2$ precipitates, and PbS precipitates last.
d) PbF_2 precipitates first, then PbS precipitates, and $Pb_3(PO_4)_2$ precipitates last.

e) $Pb_3(PO_4)_2$ precipitates first, then PbS precipitates, and PbF_2 precipitates last.

From calculate $[Pb^{2+}]$, PbS will precipitate first (when $[Pb^{2+}] > 7 \times 10^{-25} \text{ M}$), $Pb_3(PO_4)_2$ will precipitate second (when $[Pb^{2+}] > 4.7 \times 10^{-16} \text{ M}$), and PbF_2 precipitates last (when $[Pb^{2+}] > 4.0 \times 10^{-4} \text{ M}$)

5. Let s = molar solubility (mol/L) of $Al_2(CO_3)_3(s)$. Which of the following is the correct mathematical expression relating s to the K_{sp} value for $Al_2(CO_3)_3(s)$?

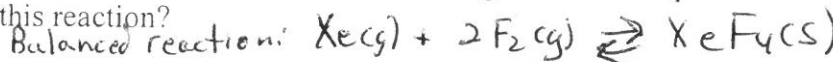


I	—	0	0
C	-s	+2s	+3s
E	—	2s	3s

$$K_{sp} = (2s)^2 (3s)^3 = 108s^5$$

6/10
19/15

6. When xenon (Xe) gas and fluorine (F₂) gas are reacted at a certain temperature, solid XeF₄ is produced. Which of the following is a correct equilibrium constant expression for this reaction?



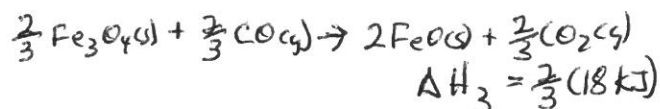
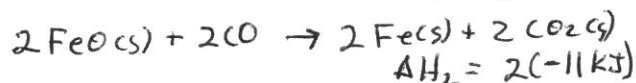
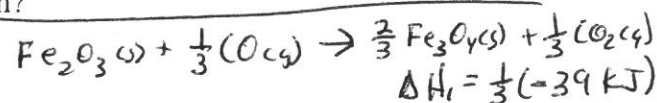
solids are not included in K expressions

- a) $\frac{1}{[\text{Xe}][\text{F}_2]^2}$ b) $\frac{[\text{Xe}][\text{F}_2]}{[\text{XeF}_4]}$ c) $\frac{[\text{XeF}_4]}{[\text{Xe}][\text{F}_2]}$
d) $\frac{[\text{XeF}_4]}{[\text{Xe}][\text{F}_2]^2}$ e) $\frac{1}{[\text{Xe}][\text{F}_2]}$ $K = \frac{1}{[\text{Xe}][\text{F}_2]^2}$

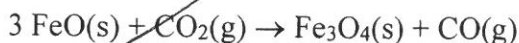
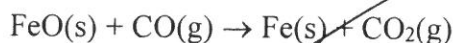
7/11
20/16

7. Consider the following reaction:
 $\text{H}_2\text{(g)} + \text{F}_2\text{(g)} \rightarrow 2\text{HF(g)}$
When product bonds are stronger than the reactant bonds for a gas phase reaction, then overall reaction must be exothermic.
Given that the HF bond is a stronger bond than either the H₂ bond or the F₂ bond, which of the following conclusions (a-d) can be drawn?

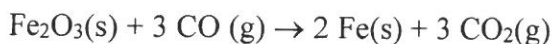
- a) The reaction must be exothermic.
b) The reaction may be exothermic.
c) The reaction may be endothermic.
d) The reaction must be endothermic.
e) Answers b and c could both be correct.



8. Given the following data:



calculate ΔH for:

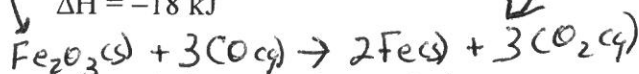


- a) -68 kJ b) 68 kJ c) -32 kJ d) 32 kJ

$$\Delta H = -11 \text{ kJ}$$

$$\Delta H = -39 \text{ kJ}$$

$$\Delta H = -18 \text{ kJ}$$



$$\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$$

$$\Delta H = ?$$

$$\Delta H = -13 \text{ kJ} - 22 \text{ kJ} + 12 \text{ kJ}$$

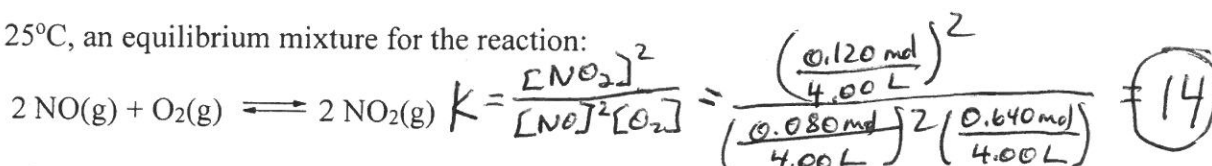
$$\Delta H = -23 \text{ kJ}$$

e) -23 kJ

Hess's Law
Problem.
See above
for initial
work

9/13
22/18

9. At 25°C, an equilibrium mixture for the reaction:



contains 0.120 mol of NO₂, 0.080 mol of NO, and 0.640 mol of O₂ in a 4.00 L bulb.
What is the value of K at 25°C for this reaction?

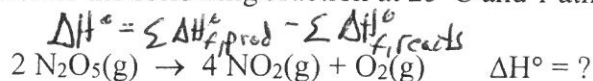
- a) 0.071 b) 14 c) 9.4 d) 3.5 e) 2.3

Form
A/B
C/D

10/1
14/23

Since PbX_2 is more soluble in acidic solution, X^- must react with H^+ . This occurs when X^- is the anion of a weak acid. Since PbY_2 is not more soluble in acidic solution, Y^- must not react with H^+ . This occurs when Y^- is the anion of a strong acid (Br^- , I^- , ClO_4^-).

10. Two salts, PbX_2 and PbY_2 , have similar solubilities in water. Experiments reveal that PbX_2 is more soluble in acidic solution than in water. However, this is not the case for the solubility of PbY_2 . Which of the following statements is true?
- ☐ a) The K_{sp} value for PbX_2 must be smaller than the K_{sp} value for PbY_2 .
- ☐ b) The K_{sp} value for PbY_2 must be smaller than the K_{sp} value for PbX_2 .
- ☐ c) Pb^{2+} is a base.
- ☒ d) The X^- anion could be F^- . Since HF is a weak acid, F^- will react with H^+ .
- ☐ e) The Y^- anion could be $C_2H_3O_2^-$. Since $HC_2H_3O_2$ is a weak acid, $C_2H_3O_2^-$ will react with H^+ . But PbY_2 did not become more soluble in acidic solution, so Y^- cannot be $C_2H_3O_2^-$.
11. Consider the following reaction at $25^\circ C$ and 1 atm:



If the standard enthalpies of formation for $N_2O_5(g)$ and $NO_2(g)$ are 11.3 kJ/mol and 33.2 kJ/mol , respectively, calculate the enthalpy change for the above reaction.

$$\Delta H^\circ = [4(33.2 \text{ kJ}) + 0] - [2(11.3 \text{ kJ})] = 110.2 \text{ kJ}$$

- a) -77.5 kJ b) 21.2 kJ c) -21.2 kJ d) 43.8 kJ e) 110.2 kJ

$$w = -P\Delta V$$

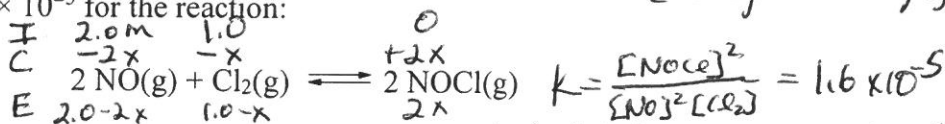
12. If the internal energy of a thermodynamic system increases by $300. \text{ J}$ while 75 J of expansion work is done, what is the value of q for this system?

$$\Delta E = q + w, \quad 300 = q - 75, \quad q = 375$$

- a) 375 J b) -375 J c) 225 J d) -225 J e) 0 J

Expansion work means $\Delta V > 0$, so system does work on surroundings (a negative sign).

13. At $35^\circ C$, $K = 1.6 \times 10^{-5}$ for the reaction:



If 2.0 mol NO and 1.0 mol Cl_2 are placed into a 1.0 L evacuated container at $35^\circ C$, what is the equilibrium concentration of $NOCl$?

$$\text{Assume } 2-2x \approx 2 \text{ and } 1-x \approx 1: \quad 1.6 \times 10^{-5} = \frac{(2x)^2}{(2-2x)^2 (1-x)} \approx \frac{4x^2}{(2.0)^2 (1.0)} = x^2$$

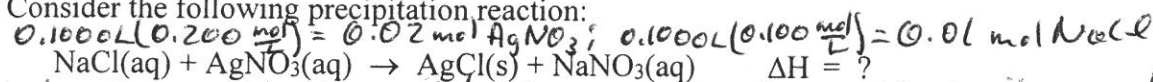
- a) $1.6 \times 10^{-5} \text{ M}$ b) $6.4 \times 10^{-5} \text{ M}$ c) 2.0 M

$$x = \sqrt{1.6 \times 10^{-5}} = 4.0 \times 10^{-3} \text{ M}; \text{ Assumptions good.}$$

- d) $8.0 \times 10^{-3} \text{ M}$ e) $4.0 \times 10^{-3} \text{ M}$

$$[NOCl]_e = 2x = 2(4.0 \times 10^{-3}) = 8.0 \times 10^{-3} \text{ M}$$

14. Consider the following precipitation reaction:



When 0.02 mol AgNO_3 reacts with 0.01 mol NaCl , $NaCl$ is limiting and 0.01 mol AgCl can be produced.

When 100.0 mL of 0.200 M AgNO_3 is mixed with a 100.0 mL of 0.100 M NaCl in a coffee-cup calorimeter, the temperature increases from $24.60^\circ C$ to $25.30^\circ C$. Determine ΔH for the above reaction. Assume the density of the solution is 1.00 g/mL and assume the heat capacity of the solution is $4.18 \text{ J/g}^\circ C$.

$$q_{\text{surroundings}} = S \times \text{mass} \times \Delta T = 4.18 \frac{\text{J}}{\text{g}^\circ C} \times 200.0 \text{ g} \times 0.70^\circ C = 585.2 \text{ J}$$

- a) -29 kJ b) -59 kJ c) 59 kJ d) -590 J e) 290 J

$$\Delta H = \frac{-585.2 \text{ J}}{0.01 \text{ mol AgCl}} = -5.85 \times 10^4 \text{ J/mol} = -59 \text{ kJ/mol}$$

Since temperature of calorimeter increased, reaction must be exothermic.

11/2
15/24

12/3
16/25

13/4
17/26

14/5
18/27

$$Q = \frac{[B_2H_6]_0 [O_2]_0^3}{[H_2O]_0^3} = \frac{(6.0 \text{ mol}) (2.0 \text{ mol})^3}{(5.0 \text{ mol})^3} = 0.384$$

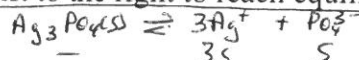
$Q > K$

15. Consider the following reaction at 25°C:

$B_2O_3(s) + 3 H_2O(g) \rightleftharpoons B_2H_6(g) + 3 O_2(g)$ $\Delta H = 2035 \text{ kJ}$, $K = 4.8 \times 10^{-2}$
Since $Q > K$, rxn shifts left to reach equilibrium (all reactant amounts increase and all product amounts decrease).
Initially, 8.0 mol B_2O_3 , 5.0 mol H_2O , 6.0 mol B_2H_6 , and 2.0 mol O_2 are mixed in a 1.0 L container. What will happen to the amount of B_2O_3 present as the reaction reaches equilibrium? B_2O_3 , like all reactants, will increase in amount as reaction shifts left to reach equilibrium.

- a) It will increase because the reaction will shift to the left to reach equilibrium.
b) It will not change because solids are not involved in the equilibrium.
c) It will not change because the reaction is already at equilibrium.
d) It will decrease because the reaction will shift to the right to reach equilibrium.

Let s = molar solubility



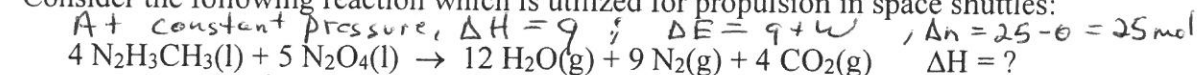
Excess $Ag_3PO_4(s)$ is added to some water. At equilibrium, the concentration of Ag^+ is $4.8 \times 10^{-5} \text{ M}$. Calculate the K_{sp} value for $Ag_3PO_4(s)$.

$K_{sp} = (3s)^3 s = 27s^4$; from problem, $[Ag^+]_e = 4.8 \times 10^{-5} \text{ M} = 3s$, $s = 1.6 \times 10^{-5} \text{ M}$

- a) 1.4×10^{-16} b) 2.3×10^{-9} c) 4.4×10^{-13} d) 6.7×10^{-10} e) 1.8×10^{-18}

$K_{sp} = 27s^4 = 27(1.6 \times 10^{-5})^4$, $K_{sp} = 1.8 \times 10^{-18}$

Consider the following reaction which is utilized for propulsion in space shuttles:



$$\omega = -P\Delta V = -RT\Delta n, \quad \omega = -8.3145 \text{ J/K} \cdot \text{mol} \times 383 \text{ K} \times 25 \text{ mol} = -7.96 \times 10^4 \text{ J}$$

At 1.00 atm and 110.°C, $\Delta E = -4674 \text{ kJ}$ for the above reaction. Determine ΔH for this reaction at 1.00 atm and 110.°C.

$$\Delta E = q + \omega, \quad -4674 \text{ kJ} = q + (-79.6 \text{ kJ}), \quad q = \Delta H = -4594 \text{ kJ}$$

- a) -4754 kJ b) -4623 kJ c) -4725 kJ d) -4594 kJ e) -4674 kJ

18. Which of the following statements (a-d) is **true** if ΔE for a system equals -95 J

($\Delta E = -95 \text{ J}$)? $\Delta E_{univ} = 0 = \Delta E_{sys} + \Delta E_{surr}$, $\Delta E_{surr} = -\Delta E_{sys}$

Since $\Delta E_{sys} = -95 \text{ J}$, $\Delta E_{surr} = -(-95) = +95 \text{ J}$

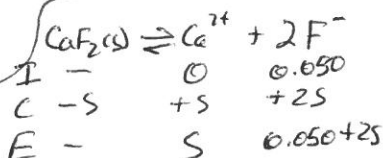
- a) The system is gaining 95 J of energy, while the surroundings are losing 95 J of energy. (system loses energy, surroundings gain energy)

- b) The system is losing 95 J of energy, while the surroundings are gaining 95 J of energy.

- c) Both the system and the surroundings are gaining 95 J of energy.

- d) Both the system and the surroundings are losing 95 J of energy.

- e) None of the above (a-d) are true.



Let s = molar solubility; $K_{sp} = [Ca^{2+}][F^-]^2 = 4.0 \times 10^{-11}$

19. How many moles of $CaF_2(s)$ will dissolve in 3.0 liters of a 0.050 M NaF solution?

Assume no volume change on addition of $CaF_2(s)$. K_{sp} for $CaF_2 = 4.0 \times 10^{-11}$.

From set-up on right, $K_{sp} = 4.0 \times 10^{-11} = s(0.050+2s)^2$

- a) $8.9 \times 10^{-9} \text{ mol}$ b) $8.0 \times 10^{-8} \text{ mol}$ c) $4.8 \times 10^{-8} \text{ mol}$

Assume $0.050+2s \approx 0.050$

- d) $2.7 \times 10^{-8} \text{ mol}$ e) 0.15 mol

$$4.0 \times 10^{-11} = s(0.050)^2$$

$$s = 1.6 \times 10^{-8} \text{ M}$$

Assumption good:

$$0.050 + s \approx 0.050$$

$$\text{mol } CaF_2 = 3.0 \text{ L} \left(\frac{1.6 \times 10^{-8} \text{ mol } CaF_2}{\text{L}} \right) = 4.8 \times 10^{-8} \text{ mol } CaF_2$$

The standard enthalpy of formation for a substance refers to the reaction of forming one mole of that substance from the elements that make up that substance all in their standard state at 1 atm and 25°C.

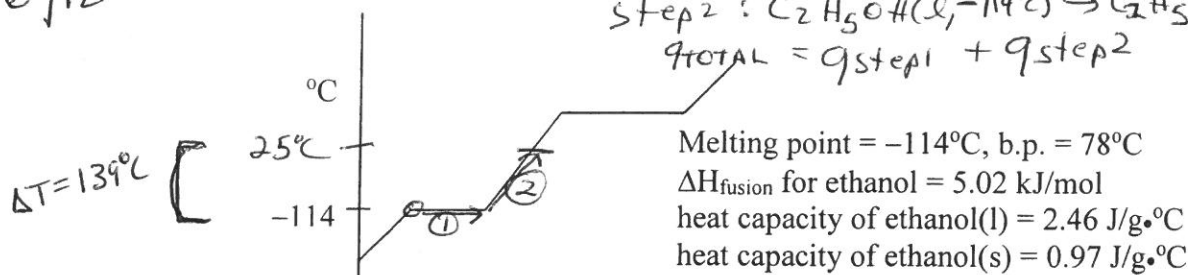
20. For how many of the following (I-IV) does the enthalpy change for the reaction equal the standard enthalpy of formation for the product in the reaction, i.e., for how many does $\Delta H_{\text{reaction}} = \Delta H_f^\circ$?

- NO I. $(1/2)\text{H}_2(\text{g}) \rightarrow \text{H}_2(\text{g})$ Don't find hydrogen as individual atoms at 1 atm + 25°C.
NO II. $\text{H}_2(\text{g}) + \text{SO}_4(\text{l}) \rightarrow \text{H}_2\text{SO}_4(\text{l})$ Don't find sulfur and oxygen as sox in standard state.
NO III. $\text{C}(\text{g}) + 4\text{H}(\text{g}) + \text{O}(\text{g}) \rightarrow \text{CH}_3\text{OH}(\text{g})$ Don't find hydrogen and oxygen as individual atoms in standard state.
NO IV. $\text{Ba}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq}) \rightarrow \text{BaCl}_2(\text{s})$ Barium standard state is $\text{Ba}(\text{s})$ and chlorine standard state is $\text{Cl}_2(\text{g})$. The correct equation for $\text{BaCl}_2(\text{s})$ is: $\text{Ba}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow \text{BaCl}_2(\text{s})$ $\Delta H = \Delta H_f^\circ(\text{BaCl}_2(\text{s}))$

- a) 0 (none) b) 1 c) 2 d) 3 e) 4 (all)

21. Consider the following heating curve for ethanol, $\text{C}_2\text{H}_5\text{OH}$:

21/15
6/12



$$q_{\text{step 1}} = 23.0 \text{ g C}_2\text{H}_5\text{OH} \left(\frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.0 \text{ g}} \right) \left(\frac{5.02 \text{ kJ}}{\text{mol C}_2\text{H}_5\text{OH}} \right) = 2.51 \text{ kJ}$$

How much energy is required to convert 23.0 g of solid ethanol at -114°C to liquid ethanol at 25°C ? The molar mass of ethanol is 46.0 g/mol .

$$q_{\text{step 2}} = 2.46 \text{ J/g}\cdot^\circ\text{C} (23.0 \text{ g}) (25 - (-114)) = 7.86 \times 10^3 \text{ J} = 7.86 \text{ kJ}$$

- a) 5.35 kJ b) 2.51 kJ c) 7.86 kJ d) 10.37 kJ e) 5.04 kJ

$$q_{\text{TOTAL}} = 2.51 \text{ kJ} + 7.86 \text{ kJ} = 10.37 \text{ kJ}$$

22. Which of the following four salts when dissolved in water will produce the largest concentration of Pb^{2+} at equilibrium? The K_{sp} to solubility relationship is shown for each salt. Also, the calculated solubility is shown. All were calculated by setting up the ICE table for each salt, then solving the problem.

PbCO_3 $K_{\text{sp}} = 1.5 \times 10^{-15} = s^2$, $s = 3.9 \times 10^{-8} \text{ mol/L}$

PbS $K_{\text{sp}} = 7.0 \times 10^{-29} = s^2$, $s = 8.4 \times 10^{-15} \text{ mol/L}$

Pb(OH)_2 $K_{\text{sp}} = 1.2 \times 10^{-15} = 4s^3$, $s = 6.7 \times 10^{-6} \text{ mol/L}$

PbCrO_4 $K_{\text{sp}} = 2.0 \times 10^{-16} = s^2$, $s = 1.4 \times 10^{-8} \text{ mol/L}$

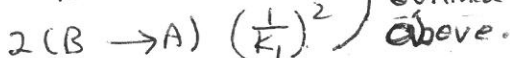
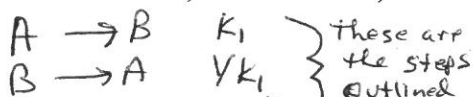
The salt with the largest solubility value will have the largest

- a) PbCO_3 b) PbS c) Pb(OH)_2 d) PbCrO_4

concentration of Pb^{2+} . This is $\text{Pb(OH)}_2(\text{s})$.

23. An initial chemical reaction is reversed and then the coefficients are all multiplied by a factor of 2. The K value for this final reaction is 0.01. What is the value of K for the initial reaction? When a reaction is reversed, $K_{\text{new}} = 1/K_{\text{old}}$. When a reaction is multiplied by some number n, $K_{\text{new}} = (K_{\text{old}})^n$.

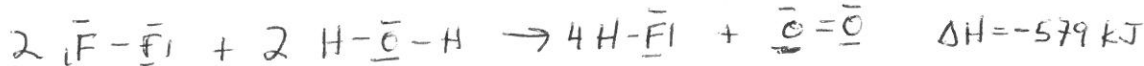
- a) 0.01 b) 100 c) 0.1 d) 10 e) 10,000



This equals 0.01

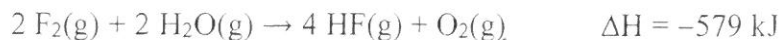
$$\left(\frac{1}{K_1} \right)^2 = 0.01, \frac{1}{K_1} = 0.1$$

$$K_1 = \frac{1}{0.1} = 10$$

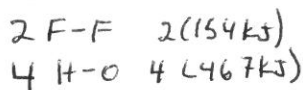
A/B
C/D

$$\Delta H_{\text{rxn}} = \Delta H_{\text{broken}} + \Delta H_{\text{formed}}$$

24. Consider the reaction:

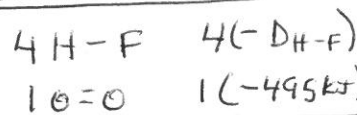
Let $D_{\text{H-F}} = \text{H-F bond energy}$

Calculate the H-F bond energy (in kJ/mol) given the following bond energies:

Bonds Broken

$$\Delta H_{\text{broken}} = 2176 \text{ kJ}$$

bond	bond energy
H-H	432 kJ/mol
O-H	467 kJ/mol
F-F	154 kJ/mol
O-O	146 kJ/mol
O=O	495 kJ/mol

Bonds Formed

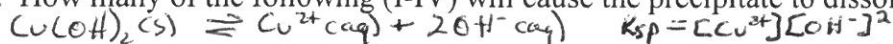
$$\Delta H_{\text{formed}} = -4D_{\text{H-F}} - 495 \text{ kJ}$$

$$\Delta H_{\text{rxn}} = \Delta H_{\text{broken}} + \Delta H_{\text{formed}} = 2176 + (-4D_{\text{H-F}} - 495) = -579 \text{ kJ}$$

- a) 314 kJ/mol b) 159 kJ/mol
- c) 565 kJ/mol
- d) 515 kJ/mol e) 2149 kJ/mol

Solving: $D_{\text{H-F}} = 565 \text{ kJ/mol}$

25. Some
- $\text{Cu}(\text{NO}_3)_2(\text{aq})$
- and
- $\text{KOH}(\text{aq})$
- are added together and a precipitate of
- $\text{Cu}(\text{OH})_2(\text{s})$
- forms. How many of the following (I-IV) will cause the precipitate to dissolve?



- NO I. Excess $\text{Cu}(\text{NO}_3)_2(\text{aq})$ is added. Added Cu^{2+} or OH^{-} causes the K_{sp} reaction to shift left causing more $\text{Cu}(\text{OH})_2(\text{s})$ precipitate to form, not less. So I is not correct.
- NO II. Excess $\text{KOH}(\text{aq})$ is added.
- Yes III. Excess $\text{NH}_3(\text{aq})$ is added. Added NH_3 removes Cu^{2+} ion as $\text{Cu}(\text{NH}_3)_2^{2+}$. As Cu^{2+} is removed, more $\text{Cu}(\text{OH})_2(\text{s})$ dissolves.
- Yes IV. Excess $\text{HNO}_3(\text{aq})$ is added. Added H^{+} from HNO_3 reacts with OH^{-} to form H_2O . As OH^{-} ion is removed, more $\text{Cu}(\text{OH})_2(\text{s})$ will dissolve to replenish the OH^{-} . So III + IV both cause the precipitate of $\text{Cu}(\text{OH})_2$ to dissolve.

- a) 0 (none) b) 1 c) 2 d) 3
- e) 4 [All will cause the $\text{Cu}(\text{OH})_2(\text{s})$ to dissolve.]

26. A bomb calorimeter has a heat capacity of
- $5.02 \text{ kJ/}^\circ\text{C}$
- and an initial temperature of
- 24.62°C
- . What is the final temperature in the bomb calorimeter if
- 1.785 g
- of benzoic acid (
- $\text{HC}_7\text{H}_5\text{O}_2$
-) is combusted in the calorimeter? The energy of combustion of benzoic acid is
- -26.42 kJ/g
- .

$$1.785 \text{ g HC}_7\text{H}_5\text{O}_2 \left(\frac{26.42 \text{ kJ}}{\text{g}} \right) = 47.16 \text{ kJ}^{\text{of heat}} \text{ added to calorimeter.}$$

Note that temp of calorimeter will increase when this heat is added.

$$47.16 \text{ kJ} = 5.02 \text{ kJ/}^\circ\text{C} \times \Delta T, \Delta T = 9.39^\circ\text{C}, T_f = 24.62^\circ\text{C} + 9.39^\circ\text{C} = 34.01^\circ\text{C}$$

27. A reaction at equilibrium shifts left to reestablish equilibrium when the volume of the container is increased. For the same reaction, the value of the equilibrium constant increases as temperature increases. Which of the following could be this reaction?
- Reactions that shift left when the container volume increases all have more reactant moles of gas than product moles of gas. These are reactions
- a) $2 \text{H}_2\text{O}_2(\text{l}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$ This reaction is exothermic. c and d.
- b) $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$ This reaction is endothermic.
- c) $3 \text{O}_2(\text{g}) \rightarrow 2 \text{O}_3(\text{g})$ This reaction is endothermic.
- d) $4 \text{Fe}(\text{s}) + 3 \text{O}_2(\text{g}) \rightarrow 2 \text{Fe}_2\text{O}_3(\text{s})$ This reaction is exothermic.

Reactions whose K value increases with a temperature increase are endothermic reactions (heat is a reactant). So reactions b and c are ones whose K value increases as T increases. Combining the two observations, reaction C fits both.

Form
A/B
C/D

CHEMISTRY 102
Hour Exam III

Set-up to this ICA-ICE problem is below.
Note that the initial concentrations of Ce^{3+} and IO_3^- have been halved. This is because the volume doubled when we added the two solutions together ($M_1V_1 = M_2V_2$).

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28. When 100.0 mL of 2.0 M $\text{Ce}(\text{NO}_3)_3$ is added to 100.0 mL of 3.0 M KIO_3 , a precipitate of $\text{Ce}(\text{IO}_3)_3(\text{s})$ forms. Calculate the equilibrium concentration of IO_3^- in this solution. K_{sp} for $\text{Ce}(\text{IO}_3)_3$ is 3.0×10^{-10} .

$$K_{\text{sp}} = 3.0 \times 10^{-10} = (0.50 + s)(3s)^3$$

a) $8.4 \times 10^{-4} \text{ M}$ b) $5.5 \times 10^{-5} \text{ M}$

Assuming $0.50 + s \approx 0.50$

d) $1.1 \times 10^{-10} \text{ M}$ e) 1.5 M

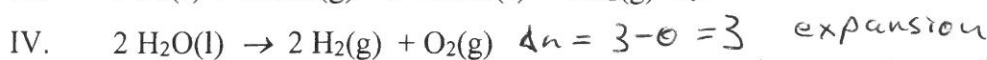
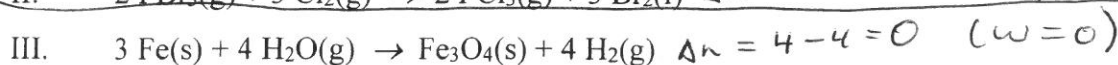
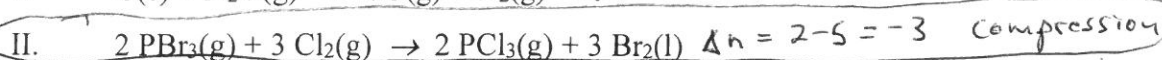
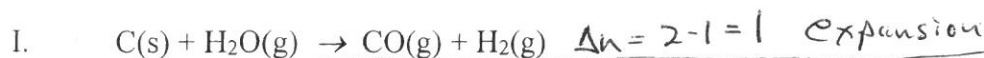
$3.0 \times 10^{-10} \approx 0.50(27s^3)$
 $s = \text{solubility} = 2.81 \times 10^{-4} \text{ M (assumption good)}$

$[\text{IO}_3^-] = 3s = 3(2.81 \times 10^{-4}) = 8.4 \times 10^{-4} \text{ M}$

	$\text{Ce}(\text{IO}_3)_3(\text{s})$	$\rightleftharpoons$	Ce^{3+}	$+ 3\text{IO}_3^-$
I			1.0 M	1.5 M
C			-0.50	-1.5
A			0.50	0
I			0.50	0
C - s			+s	+3s
E			0.50 + s	3s

29/29
29/29

29. Consider the following four reactions at constant pressure:



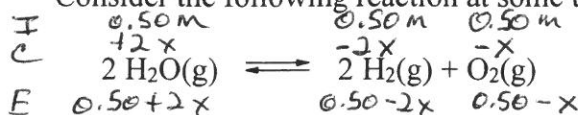
$w = -P\Delta V$. Work is a positive value when ΔV is negative. This occurs in a compression where the moles of gas decrease as reactants are converted to products.

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

e) 4 (All of these reactions have a positive value for the PV work.)

Note that in a compression, the surroundings do work on the system which is a positive value for w . Only reaction II is a compression.

30. Consider the following reaction at some temperature:



when rate reverse > rate forward, reaction shifts left to reach equilibrium.

Initially, 1.0 mol of H_2 , 1.0 mol of O_2 , and 1.0 mol of H_2O are placed in a 2.0 L container. At these conditions, the rate of the reverse reaction is greater than the rate of the forward reaction. Which of the following statements (a-d) regarding this reaction is true once it reaches equilibrium?

- a) $K > 1$. $Q = \frac{(0.50)^2(0.50)}{(0.50)^2} = 0.50$; as reaction shifts left to reach equilibrium, the products get smaller than 0.50 M, while the reactants get bigger. This means $K < 1$.
- b) At equilibrium, $[\text{H}_2\text{O}] = [\text{H}_2]$. From ICE Table, these concentrations can't be equal at equilibrium.
- c) At equilibrium, $[\text{H}_2\text{O}] > 0.50 \text{ M}$. $[\text{H}_2\text{O}]_e = 0.50 + x$, so $[\text{H}_2\text{O}]_e > 0.50 \text{ M}$.
- d) At equilibrium, $[\text{H}_2] = [\text{O}_2]$. From ICE Table, these two concentrations are not equal ($[\text{H}_2]_e = 0.50 - 2x$ and $[\text{O}_2]_e = 0.50 - x$).
- e) None of these statements (a-d) are true.

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