How many of the following five silver salts are more soluble in acidic (H⁺) solution than 1. in pure water?

> AgCl AgI Ag_2CO_3 AgBr AgClO₄

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

- e) 5 (All are more soluble in acidic solution than in pure water.)
- 2. Calculate the average carbon-fluorine bond energy using the following information:



Bond	Bond Energy (kJ/mol)
С–Н	413
C=C	614
C–C	347
H–H	432
F–F	154

The average C–F bond energy is equal to:

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

- 3. Which of the following statements is **false**?
 - a) The energy of the universe is constant.
 - b) Heat is a path function.
 - c) The internal energy of a system is the sum of its potential and kinetic energies.
 - d) When a gas expands against a constant external pressure, the surroundings are doing work on the system.
- 4. Which of the following statements (a-c) correctly relates chemical potential energy and bond strength?

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- 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.

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

$\Delta H_{vap} = 20.00 \text{ kJ/mol}$	specific heat capacity of solid = $3.00 \text{ J/g} \cdot ^{\circ}\text{C}$
$\Delta H_{fus} = 5.00 \text{ kJ/mol}$	specific heat capacity of liquid = $2.50 \text{ J/g} \cdot ^{\circ}\text{C}$
boiling point = 75.0°C	molar mass = 100.0 g/mol
melting point = -15.0° C	

5. 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?

a) 78.9 min b) 7.89 min c) 3.33 min d) 4.67 min e) 3.56 min

6. 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).

a) $2.00 \text{ J/g} \cdot ^{\circ}\text{C}$	b) 1.00 J/g • °C	c) $1.50 \text{ J/g} \cdot ^{\circ}\text{C}$
d) 0.500 J/g•°C	e) 1.75 J/g.°C	

- 7. When H₂(g) and Cl₂(g) are reacted in a 4.0 L container, HCl(g) is produced in an exothermic reaction. Which of the following changes (a-d) would <u>not</u> cause the reaction to shift right to reestablish equilibrium?
 - a) Some HCl(g) is removed.
 - b) The temperature is decreased.
 - c) More $H_2(g)$ is added.
 - d) The volume of the container is decreased to 2.0 L.
 - e) All of the changes (a-d) would cause the reaction to shift right.
- 8. Consider the following reaction at $T = 25^{\circ}C$:

 $H_2(g) + I_2(g) \rightleftharpoons 2 HI(g)$ K = 1000

Which of the following statements is **false**?

- a) The value of the equilibrium constant depends on the amounts of reactants and products that are mixed together initially.
- b) For this reaction, $K_p = K$.
- c) At equilibrium, the rate of the forward reaction equals the rate of the reverse reaction.
- d) At equilibrium, $[HI]^2 > [H_2] \times [I_2]$.
- e) At equilibrium, the concentration of reactants and products are constant.

- 9. ΔH° and ΔE° are approximately equal to each other in all of the following constant pressure processes **except** for:
 - a) $COCl_2(g) \rightarrow CO(g) + Cl_2(g)$ b) $H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Na^+(aq) + Cl^-(aq)$ c) $H_2(g) + I_2(g) \rightarrow 2 HI(g)$ d) $CO(g) + H_2O(g) \rightarrow H_2(g) + CO_2(g)$ e) $Hg(s) \rightarrow Hg(l)$
- 10. Consider the solubility of solid Cd(OH)₂ in the various solutions (I-III) below:
 - I. Solubility of $Cd(OH)_2(s)$ in 1.0 *M* HNO₃
 - II. Solubility of Cd(OH)₂(s) in pure water
 - III. Solubility of $Cd(OH)_2(s)$ in 1.0 *M* NaOH

Rank the solubility of AgOH in solutions I-III from lowest molar solubility to highest molar solubility.

a) $I < II < III$ b) $I < III <$	< II c) II $<$ I $<$ III
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- d) II < III < I e) III < II < I f) III < II < II
- 11. Which of the following statements is **false**?
 - a) In a bomb calorimeter, ΔE can be determined.
 - b) When an endothermic reaction is performed in a calorimeter, the temperature of the calorimeter increases.
 - c) An element in its standard state has a standard enthalpy of formation equal to zero $(\Delta H_f^\circ = 0)$.
 - d) Hess's law is based on the fact that enthalpy is a state function.
 - e) In a coffee cup calorimeter, the heat released or gained at constant pressure (q_p) can be determined.
- 12. Consider the following endothermic reaction:

 $N_2(g) + O_2(g) \rightarrow 2 NO(g)$

Which of the following statements is **false** concerning this reaction?

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

13. Consider the following reaction at some constant temperature:

 $2 \operatorname{CO}_2(g) \rightleftharpoons 2 \operatorname{CO}(g) + \operatorname{O}_2(g)$ $K = 2.0 \times 10^{-6}$

In an experiment, 4.0 mol of CO₂ and 2.0 mol of CO are each placed into a 1.0 L rigid container. Calculate the concentration of O_2 ([O₂]_e = ?) once the reaction has reached equilibrium.

a) $5.0 \times 10^{-7} M$ b) $1.0 \times 10^{-6} M$ c) $2.0 \times 10^{-6} M$ d) $4.0 \times 10^{-6} M$ e) $8.0 \times 10^{-6} M$

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

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

a) 1 b) 2 c) 3 d) 4 e) 5 (All are intensive properties.)

15. Consider the following reaction at some constant temperature:

 $H_2(g) + I_2(g) \rightleftharpoons 2 HI(g) \qquad K_p = 0.010$

In an experiment, $H_2(g)$ at 1.00 atm, $I_2(g)$ at 1.00 atm, and 1.00 atm of HI are mixed together initially. At equilibrium, what is the partial pressure of HI(g)?

a) 0.57 atm b) 0.43 atm c) 0.14 atm d) 0.25 atm e) 0.50 atm

16. Consider the following reaction:

 $4 H_2O(g) + 3 Fe(s) \rightleftharpoons 4 H_2(g) + Fe_3O_4(s)$

What is the correct equilibrium constant expression for this reaction?

a)
$$K = \frac{[H_2O]^4[Fe]^3}{[H_2]^4[Fe_3O_4]}$$
 b) $K = \frac{[H_2][Fe_3O_4]}{[H_2O][Fe]}$ c) $K = \frac{[H_2]}{[H_2O]}$

d)
$$K = \frac{[H_2]^4[Fe_3O_4]}{[H_2O]^4[Fe]^3}$$
 e) $K = \frac{[H_2]^4}{[H_2O]^4}$

17. 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:

 $HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq) \qquad K = ?$

Given that acetic acid $(HC_2H_3O_2)$ is a weak electrolyte (weak acid), which of the following is most likely the K value for the above equilibrium reaction?

a)
$$7.2 \times 10^{27}$$
 b) 6.1×10^{14} c) 5.4×10^{12} d) 1.5 e) 1.8×10^{-5}

18. Given:

$2 \operatorname{Cu}_2 O(s) + O_2(g) \rightarrow 4 \operatorname{Cu}O(s)$	$\Delta H^{\circ} = -288 \text{ kJ}$
$Cu_2O(s) \rightarrow CuO(s) + Cu(s)$	$\Delta H^{\circ} = 11 \text{ kJ}$

Calculate the standard enthalpy of formation (ΔH_f°) for CuO(s).

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

19. Consider the following reaction at some constant temperature:

$$2 O_3(g) \rightleftharpoons 3 O_2(g) \qquad K = ?$$

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.

- a) 98 b) 490 c) 340 d) 3.4×10^4 e) 0.98
- 20. Excess PbCl₂(s) is added to 100.0 mL of water. Calculate the concentration of Pb²⁺ in solution ([Pb²⁺]_e = ?). K_{sp} for PbCl₂ = 1.6×10^{-5} .
 - a) 0.016 M b) 0.032 M c) 0.064 M d) 0.10 M e) $2.0 \times 10^{-3} M$

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

 $N_2O_4(g) \rightleftharpoons 2 NO_2(g)$ K = 8.0

An equilibrium mixture contains 6.0 mol of $N_2O_4(g)$ in a 3.0 L container. How many moles of $NO_2(g)$ are present in this equilibrium mixture?

- a) 2.0 mol
 b) 4.0 mol
 c) 6.0 mol

 d) 8.0 mol
 e) 12 mol
- 22. The solubility of Ag₃PO₄(s) in 0.20 *M* AgNO₃ is 2.2×10^{-16} mol/L. Calculate the K_{sp} value for Ag₃PO₄.
 - 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}

23. Which of the following statements is **false**?

- a) The internal energy of a system decreases when more work is done by the system than heat is flowing into the system.
- b) At constant volume, the internal energy change for a system is equal to the amount of heat flow ($\Delta E = q_v$).
- c) The internal energy of a system increases when work is done on the system and heat is flowing into the system.
- d) For the process of boiling water, the sign of q is negative.
- 24. The combustion of benzoic acid (C_6H_5COOH , 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.

a)	1000 kJ/°C	b) 10.0 kJ/°C	c) 1220 kJ/°C
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d) $0.379 \text{ kJ/}^{\circ}\text{C}$ e) $8.20 \text{ kJ/}^{\circ}\text{C}$

25. Consider the following solutions prepared in two separate beakers:

Beaker 1: 100.0 mL of $3.0 \times 10^{-4} M$ Pb(NO₃)₂(aq) mixed with 100.0 mL of $2.0 \times 10^{-4} M$ KI(aq); K_{sp} for PbI₂ = 1×10^{-8} .

Beaker 2: 100.0 mL of $3.0 \times 10^{-4} M$ Pb(NO₃)₂(aq) mixed with 100.0 mL of $2.0 \times 10^{-4} M$ K₂SO₄(aq); K_{sp} for PbSO₄ = 2.0×10^{-8} .

In which of the beakers will a precipitate form?

- a) PbI₂(s) will form in beaker 1 and PbSO₄(s) will form in beaker 2.
- b) PbI₂(s) will form in beaker 1 and no precipitate will form in beaker 2.
- c) No precipitate will form in beaker 1 and PbSO₄(s) will form in beaker 2.
- d) No precipitate will form in either beaker.
- 26. Which of the following salts has the **largest** molar solubility in pure water?
 - a) FeCO₃, $K_{sp} = 2.1 \times 10^{-11}$ b) Ag₂CrO₄, $K_{sp} = 9.0 \times 10^{-12}$
 - c) CaF₂, $K_{sp} = 4.0 \times 10^{-11}$ d) AgCl, $K_{sp} = 1.6 \times 10^{-10}$
- 27. Consider the reaction:

 $2 \text{ CH}_3\text{OH}(1) + 3 \text{ O}_2(g) \rightarrow 2 \text{ CO}_2(g) + 4 \text{ H}_2\text{O}(l)$ $\Delta \text{H}_{rxn}^\circ = -1452 \text{ kJ}$

for which ΔH_f° for $H_2O(l) = -286 \text{ kJ/mol}$ and ΔH_f° for $CO_2(g) = -393 \text{ kJ/mol}$. Calculate the standard enthalpy of formation for methanol.

- a) 239 kJ/mol b) -478 kJ/mol c) 478 kJ/mol d) -239 kJ/mol e) 0 kJ/mol
- 28. Calculate the molar solubility of $Ba_3(PO_4)_2(s)$ in 0.10 *M* K₃PO₄. K_{sp} for $Ba_3(PO_4)_2 = 6.0 \times 10^{-39}$.
 - a) $8.9 \times 10^{-9} \text{ mol/L}$ b) $1.1 \times 10^{-13} \text{ mol/L}$ c) $2.8 \times 10^{-13} \text{ mol/L}$
 - d) $7.7 \times 10^{-20} \text{ mol/L}$ e) $1.1 \times 10^{-9} \text{ mol/L}$

29. 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 J/g • °C, calculate the final temperature of the reaction mixture.

 $HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq) \qquad \Delta H = -56 \text{ kJ/mol}$

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

- 30. When 50.0 mL of 2.0 *M* AgNO₃(aq) is mixed with 50.0 mL of 3.0 *M* Na₂CrO₄(aq), a precipitate of Ag₂CrO₄(s) forms (K_{sp} for Ag₂CrO₄ = 9.0×10^{-12}). Once equilibrium has been established, calculate the Ag⁺ concentration ([Ag⁺]_e = ?).
 - a) $1.3 \times 10^{-4} M$ b) $3.0 \times 10^{-6} M$ c) $2.6 \times 10^{-4} M$
 - d) 1.0 M e) $1.5 \times 10^{-6} M$