Name ______KEY______________ Sec ________ Score ______/10
Chemistry 1B, Fall 2015 Quiz #2A

You must show your work for full credit.

Exp 1 Question

1. Solutions (at moderate concentrations) containing which of the following compounds is expected to have a low conductivity reading? (1pt)

a) HNO_3	b) NH_3
c) $Co(NO_3)_2$	d) K_2CO_3

Net ionic equations

2. Write a **net ionic equation** for the reaction of the cobalt (II) hydroxide, with hydrochloric acid. Don't forget state symbols (1pt).
reaction: $Co(OH)_2(aq) + 2HCl(aq) \leftrightarrow 2H_2O(l) + CoCl(aq)$
all ion reaction: $Co^{2+}(aq) + 2OH^-(aq) + 2H^+(aq) + 2Cl^-(aq) \leftrightarrow Co^{2+}(aq) + 2Cl^-(aq) + 2H_2O(l)$
net ionic reaction: $2OH^-(aq) + 2H^+(aq) \leftrightarrow 2H_2O(l)$ *(also o.k. with the coefficient of 2 removed)*

3. Write a **net ionic equation** for the reaction of solid silver carbonate, Ag_2CO_3, with aqueous nitric acid. Remember how nitric acid exists in solution. Don't forget state symbols (1pt).
reaction: $Ag_2CO_3(s) + 2HNO_3(aq) \leftrightarrow 2AgNO_3(aq) + CO_2(g) + H_2O(l)$
all ion reaction: $Ag_2CO_3(s) + 2H^+(aq) + 2NO_3^-(aq) \leftrightarrow 2Ag^+(aq) + 2NO_3^-(aq) + CO_2(g) + H_2O(l)$
net ionic reaction: $Ag_2CO_3(s) + 2H^+(aq) \leftrightarrow 2Ag^+(aq) + CO_2(g) + H_2O(l)$

LeChâtelier's problem

4. Consider the equilibrium reaction below and answer the following questions a-d (4pts):

$$H_2(g) + I_2(s) \leftrightarrow 2HI(g) \quad \Delta H_{rxn} = +26 \text{ kJ/mol}$$

a) If this mixture is transferred from a 2 L flask to a 8 L flask, in which direction will the equilibrium shift? Assume you still have each compound present in the same state. Explain.
The volume increases so the side with more moles of gas is favored. This is the products.
(2 mol HI vs. 1 mol H_2 *and* I_2 *doesn't count because it is a solid)*

b) Water is added which removes HI (g). Which direction does the equilibrium shift?
This removes a product. That shifts the reaction toward the products.

c) The temperature of the reaction mixture decreases. Which direction does the equilibrium shift?
Because $\Delta H>0$*, we can consider heat a reactant. Increasing T decreases the products.*

d) More I_2 (s) is added. How does the equilibrium change?
No change. Because I_2 *is a solid it is not in the equilibrium equation.*

One more on back...

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Calculating an equilibrium concentration when we know K

5. Consider the following equation:

$$2H_2S(g) \leftrightarrow 2H_2(g) + S_2(g) \quad K_C = 1.91 \times 10^{-8} \text{ (at 1000 K)}$$

If the initial concentration of $H_2S(g)$ was 0.060 M and the other species were at zero, what is the equilibrium concentration of the product, $H_2(g)$? (3pts) You must show your work for full credit. If you make any simplifying assumptions, show and validate them.

In case needed, the quadratic equation for $ax^2 + bx + c = 0$ is $x = \frac{-b \pm \sqrt{b^2-4ac}}{2a}$

The equilibrium equation is: $K_C = 1.91 \times 10^{-8} = [H_2]^2[S_2]/[H_2S]^2$

Because we are only given an initial concentration, we need to set up an ICE table

	$2H_2S(g) \leftrightarrow$	$2H_2(g) +$	$S_2(g)$
initial	*0.060 M*	*0*	*0*
change	*-2x*	*+2x*	*+x*
equil.	*0.060 – 2x*	*2x*	*x*

Putting these into the equation, we get: $1.91 \times 10^{-8} = (2x)^2x/(0.060 - 2x)^2$ *and we see we cannot solve this without a simplifying assumption. Because* K_C *is a small number, we can expect that x will be small and can make the assumption that* $0.060 >> 2x$*, replacing* $0.060 - 2x$ *with 0.060.*

Now, we get $1.91 \times 10^{-8} = (2x)^2x/(0.060)^2$

Simplifying that, we get $1.91 \times 10^{-8} = 4x^3/(0.0036)$ *and* $x^3 = 1.91 \times 10^{-8}(0.0036)/4 = 1.719 \times 10^{-11}$

or $x = (1.719 \times 10^{-11})^{1/3} = 2.58 \times 10^{-4}$

$[H_2] = 2x = 2(2.58 \times 10^{-4}) = 5.2 \times 10^{-4}$ *M*

Our assumption is valid because $2x << 0.060$ *(can write as 0.060 vs. 0.0005 so that the equilibrium* $[H_2S]$ *would barely be affected or 0.0005/0.060 = 1% change which is less than 5%)*