Name _____ Key _____   Name _____
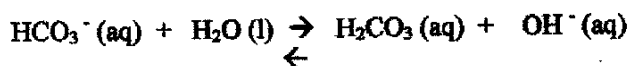(print name)   (sign name)   Please **show all work** for full credit.

1. For the following reaction, give the equilibrium constant expression as concentration in molarities in the normal format. (5 pts)

HCO_3^- (aq) + H_2O (l) $\rightleftarrows$ H_2CO_3 (aq) + OH^- (aq)

$$K_{eq} = \frac{[H_2CO_3][OH^-]}{[HCO_3^-]}$$

$$K_{eq} = \frac{product}{reactants}$$

you always leave off solids + liquids

2. For the reaction I_2(g) + Cl_2 (g) $\rightleftarrows$ 2 I Cl (g), K_{eq} = 4.2 x 10^{-7} (I made up these numbers.)

Initially you have [I_2 (g)] = 0.125 M and [Cl_2 (g)] = 0.250 M. What is the equilibrium concentration of the I Cl (g) in molarity ? (I am not looking for the final answer. Just set up the problem because you do not have enough time to actually complete the algebra.) I have started filling out the ICE table so that you can complete the problem. (8.5 pts total, one ½ point can be wrong with no penalty bc quiz has too many pts)

ratio is 1 of I_2 + Cl_2 + 2 of ICl

a. Fill out the table shown below. (½ pt per table block)

½ pt per blank

I_2 Cl_2 $\rightarrow$ 2 I Cl (set)

4.0 pts off 1st blank

X = lowest ratio

	[I_2]	[Cl_2]	[I Cl]
initial	0.125 m	0.250 m	0
change	- X	- X	+ 2x
equilibrium	0.125 - X	0.250 - X	2x

b. Write out the equilibrium constant expression for the equation using your equilibrium values. (2 pts)

$$K_{eq} = \frac{[ICl]^2}{[I_2][Cl_2]} = \frac{(2x)^2}{(0.125-x)(0.250-x)}$$

½ pt

½ pt

1 pt

c. Write out the equilibrium constant expression for the reaction using your equilibrium values with your simplifying approximation. (2 pts)

$$K_{eq} = \frac{(2x)^2}{(0.125)(0.250)} = \frac{4.2}{x} \times 10^{-7}$$

simplifying approximation

$x << 0.125$
& $x << 0.250$

$$4x^2 = (4.2 \times 10^{-7})(0.125)(0.250)$$

$$x = \sqrt{(4.2 \times 10^{-7})(0.125)(0.250)/4} = 5.73 \times 10^{-5}$$

1 pt

3. Give at least one strong acid and one weak acid. I would like the molecular formulas for the acids. (2 pts, 1 pt per blank)

Strong acid ___ HCl, HBr, HI, HNO_3, H_2SO_4

Weak acid ___ HF, CH_3COH, HCN, H_3PO_4

4. If I make up a 0.073 M solution of a **strong base** NaOH, what is the pH of this solution? (pH = - log [H⁺], pH + pOH = 14, [H⁺][OH⁻] = 1.0 x 10⁻¹⁴) (6 pts) use either method

$$NaOH \rightarrow Na^+ + OH^- \text{ (complete dissociation)}$$
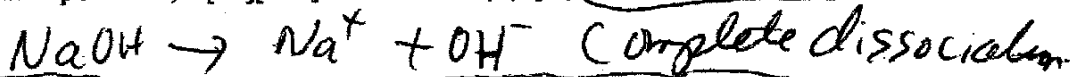
$[OH^-] = 0.073$ because 1:1 ratio   method 1

$pOH = -log(0.073) = 1.13$, $pH = 14 - pOH$

3 pts

$pH = 14 - 1.13 = 12.86$

or $[H^+][OH^-] = 1.0 \times 10^{-14}$ & $[H^+] = \frac{1.0 \times 10^{-14}}{[OH^-]}$ method 2

$$[H^+] = \frac{1.0 \times 10^{-14}}{0.073} = 1.37 \times 10^{-13}$$
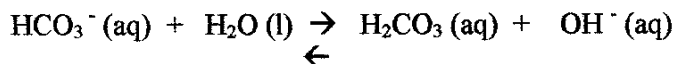
$$pH = -log(1.37 \times 10^{-13}) = 12.86$$

3 pts

3 pts

Name _____   Name _____
(print name)   (sign name)   Please **show all work** for full credit.

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2   For the reaction I_2(g) + Cl_2 (g) $\rightleftarrows$ 2 I Cl (g),   Keq = 4.2 x 10^{-7} (I made up these numbers.)

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a. Fill out the table shown below. (½ pt per table block)

I_2 Cl_2 $\rightarrow$ 2 I Cl
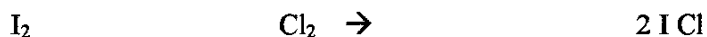
	[I_2]	[Cl_2]	[I Cl]
initial			
change			
equilibrium			

b. Write out the equilibrium constant expression for the equation using your equilibrium values. (2 pts)

c. Write out the equilibrium constant expression for the reaction using your equilibrium values with your simplifying approximation. (2 pts)

3 Give at least one strong acid and one weak acid.   I would like the molecular formulas for the acids. (2 pts, 1 pt per blank)

Strong acid _____

Weak acid _____

4. If I make up a 0.073 M solution of a **strong base** NaOH, what is the pH of this solution ?
$(pH = -\log [H^+],\ \ pH + pOH = 14,\ \ [H^+][OH^-] = 1.0 \times 10^{-14}\ )$ (6 pts)