Given this equation:

$$H_2 + I_2 \Longleftrightarrow 2\ HI$$

Calculate all three equilibrium concentrations when $[H_2]_o = [I_2]_o = 0.200$ M and $K_c = 64.0$.

The solution technique involves the use of an ICEbox.

Initial	0.200	0.200	0
Change			
Equilibrium			

H_2 + I_2 <===> 2 HI

Now for the change row. This is the one that causes the most difficulty in understanding:

	$[H_2]$	$[I_2]$	[HI]
Initial	0.200	0.200	0
Change	- x	- x	+ 2x
Equilibrium			

The minus sign comes from the fact that the H_2 and I_2 amounts are going to go down as the reaction proceeds.

The positive signifies that more HI is being made as the reaction proceeds on its way to equilibrium.

The two is important. HI is being made twice as fast as either H_2 or I_2 are being used up.

$$H_2 + I_2 \Longleftrightarrow 2\ HI$$

Calculate all three equilibrium concentrations when $[H_2]_o = [I_2]_o = 0.200$ M and $K_c = 64.0$.

	$[H_2]$	$[I_2]$	$[HI]$
Initial	0.200	0.200	0
Change	- x	- x	+ 2x
Equilibrium	0.200 - x	0.200 - x	2x

The equilibrium expression is:

$$K_c = [HI]^2 / ([H_2]\ [I_2])$$

Plugging values into the expression gives:

$$64.0 = (2x)^2 / ((0.200 - x)\ (0.200 - x))$$

Both sides are perfect squares (done so on purpose), so we square root both sides to get:

$$8.00 = (2x) / (0.200 - x)$$

From there, the solution should be easy and results in $x = 0.160$ M.

This is not the end of the solution since the question asked for the equilibrium concentrations, so:

$[H_2] = 0.200 - 0.160 = 0.040$ M
$[I_2] = 0.200 - 0.160 = 0.040$ M
$[HI] = 2\ (0.160) = 0.320$ M

Now for a second example. This example will involve the use of the quadratic formula. Given this equation:

$$PCl_3 + Cl_2 <===> PCl_5$$

Calculate all three equilibrium concentrations when K_c = 16.0 and $[PCl_5]_o$ = 1.00 M.

Here is the completed ICEbox:

	$[PCl_3]$	$[Cl_2]$	$[PCl_5]$
Initial	0	0	1.00
Change	+ x	+ x	- x
Equilibrium	x	x	1.00 - x

The equilibrium expression is:

$$K_c = [PCl_5] / ([PCl_3] [Cl_2])$$

Substituting gives:

$$16.0 = 1.00 - x / (x \text{ times } x)$$

After suitable manipulation (which you can perform yourself), we arrive at this quadratic equation in standard form:

$$16x^2 + x - 1 = 0$$

Using the quadratic formula, which is $x = (- b \pm \text{square root}[b^2 - 4ac]) / 2a$, we obtain:

$$x = (- 1 + -\text{square root}[1^2 - (4)(16)(-1)]) / 32$$

After suitable calculations, we find x = 0.221.

Please notice that the negative root was dropped, because negative b turned out to be negative one. The answer obtained in this type of problem CANNOT be negative.

Why?

Because we are dealing with the amount of a physical substance in mol / L. Amounts of substances are always represented with positive numbers. An amount of a substance with physical reality cannot be represented with negative numbers.

$$COCl_2 \Longleftrightarrow CO + Cl_2$$

Calculate all three equilibrium concentrations when $K_c = 0.680$ with $[CO]_o = 0.500$ and $[Cl_2]_o = 1.00$ M.

Here is the completed ICEbox:

	$[COCl_2]$	$[CO]$	$[Cl_2]$
Initial	0	0.500	1.00
Change	+ x	- x	- x
Equilibrium	x	0.500 - x	1.00 - x

The equilibrium expression is:

$$K_c = ([CO]\ [Cl_2]) \ / \ [COCl_2]$$

Substituting into the expression gives:

$$0.680 = ((0.5 - x)\ (1 - x)) \ / \ x$$

After some manipulation (left to the student), we arrive at this quadratic equation, in standard form:

$$x^2 - 2.18x + 0.5 = 0$$

Using the quadratic formula, we have this to start:

$$x = (2.18 \pm \text{square root}[(2.18)^2 - (4)(1)(0.5)]) / 2$$

After some manipulation (left to the student), we arrive at:

$$(2.18 \pm 1.66) / 2$$

Both roots yield positive values, so how do we pick the correct one?

The answer lies in the fact that x is not the final answer, whereas (0.5 - x) is. It is the term (0.5 - x) which must be positive.

So the root of 1.92 is rejected in favor of the 0.26 value

10 When 0.20 mol of hydrogen gas and 0.15 mol of iodine gas are heated at 723 K until equilibrium is established, the equilibrium mixture is found to contain 0.26 mol of hydrogen iodide.

The equation for the reaction is as follows.

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

What is the correct expression for the equilibrium constant K_c?

A $\dfrac{2 \times 0.26}{0.20 \times 0.15}$

hint: do the ice diagram...lol

B $\dfrac{(2 \times 0.26)^2}{0.20 \times 0.15}$

C $\dfrac{(0.26)^2}{0.07 \times 0.02}$

D $\dfrac{(0.26)^2}{0.13 \times 0.13}$

10 When vanadium(II) compounds are dissolved in water, the following equilibrium is established.

$$V^{2+} + H_2O \rightleftharpoons V^{3+} + \frac{1}{2}H_2 + OH^-$$

What would alter the composition of the equilibrium mixture in favour of the V^{2+} ions?

A adding an acid

B adding a reagent that selectively precipitates V^{3+} ions

C allowing the hydrogen to escape as it forms

D making the solution more alkaline

10 The dissociation of dinitrogen tetraoxide into nitrogen dioxide is represented by the equation below.

$$N_2O_4(g) \rightleftharpoons 2NO_2(g); \qquad \Delta H^{\ominus} = +57\,kJ\,mol^{-1}$$

If the temperature of an equilibrium mixture of the gases is increased at constant pressure, will the volume of the mixture increase or decrease and why?

A The volume will increase, but only because of a shift of equilibrium towards the right.

B The volume will increase, both because of a shift of equilibrium towards the right and also because of thermal expansion.

C The volume will stay the same, because any thermal expansion could be exactly counteracted by a shift of equilibrium towards the left.

D The volume will decrease, because a shift of equilibrium towards the left would more than counteract any thermal expansion.

2 Alcohols and esters are important organic compounds which are widely used as solvents.

Esters such as ethyl ethanoate can be formed by reacting carboxylic acids with alcohols.

$$CH_3CO_2H + C_2H_5OH \rightleftharpoons CH_3CO_2C_2H_5 + H_2O$$

This reaction is an example of a dynamic equilibrium.

(a) Explain what is meant by the term *dynamic equilibrium*.

....

....[1]

(b) Write the expression for the equilibrium constant for this reaction, K_c.

[1]

[1]

(c) For this equilibrium, the value of K_c is 4.0 at 298 K.
A mixture containing 0.5 mol of ethanoic acid, 0.5 mol ethanol, 0.1 mol ethyl ethanoate and 0.1 mol water was set up and allowed to come to equilibrium at 298 K. The final volume of solution was V dm^3.

Calculate the amount, in moles, of each substance present at equilibrium.

2 (a) rate of forward reaction equals rate of backward reaction *or* equilibrium concentrations remain constant while reaction is occurring [1] **[1]**

(b) $K_c = \dfrac{[CH_3CO_2C_2H_5][H_2O]}{[CH_3CO_2H][C_2H_5OH]}$ [1] **[1]**

(c) $CH_3CO_2H + C_2H_5OH \rightleftharpoons CH_3CO_2C_2H_5 + H_2O$

initial moles	0.5	0.5	0.1	0.1	
equil. moles	$(0.5 - x)$	$(0.5 - x)$	$(0.1 + x)$	$(0.1 + x)$	[1]
equil. concn./ mol dm^{-3}	$\dfrac{(0.5-x)}{V}$	$\dfrac{(0.5-x)}{V}$	$\dfrac{(0.1+x)}{V}$	$\dfrac{(0.1+x)}{V}$	

$K_c = \dfrac{(0.1+x)^2}{(0.5-x)^2} = 4$ [1]

gives $x = 0.3$ [1]

$n(CH_3CO_2H) = n(C_2H_5OH) = 0.2$ **and**

$n(CH_3CO_2C_2H_5) = n(H_2O) = 0.4$ [1]

allow ecf on wrong equil. moles subject to $x < 0.5$ **[4]**

NO is also formed when nitrosyl chloride, $NOCl$, dissociates according to the following equation.

$$2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$$

Different amounts of the three gases were placed in a closed container and allowed to come to equilibrium at 230 °C. The experiment was repeated at 465 °C.

The equilibrium concentrations of the three gases at each temperature are given in the table below.

	concentration / mol dm^{-3}		
temperature /°C	$NOCl$	NO	Cl_2
230	2.33×10^{-3}	1.46×10^{-3}	1.15×10^{-2}
465	3.68×10^{-4}	7.63×10^{-3}	2.14×10^{-4}

(c) (i) Write the expression for the equilibrium constant, K_c, for this reaction. Give the units.

(ii) Calculate the value of K_c at each of the temperatures given.

230 °C

465 °C

(iii) Is the forward reaction endothermic or exothermic? Explain your answer.

....

....

[5]

(d) The temperature of the equilibrium was then altered so that the equilibrium concentrations of NOC*l* and NO were the same as each other.

What will be the effect on the equilibrium concentration of NOC*l* when the following changes are carried out on this new equilibrium? In each case, explain your answer.

(i) The pressure of the system is halved at constant temperature.

....

....

(ii) A mixture of NOC*l*(g) and NO(g) containing equal numbers of moles of each gas is introduced into the container at constant temperature.

....

....

[4]

(c) (i) $K_c = \dfrac{[NO]^2[Cl_2]}{[NOCl]^2}$ (1)

units are mol dm^{-3} (1)

(ii) at 230 °C $K_c = \dfrac{(1.46\times10^{-3})^2\times1.15\times10^{-2}}{(2.33\times10^{-3})^2}$

$= 4.5 \times 10^{-3}$ mol dm^{-3} (1)

at 465 °C $K_c = \dfrac{(7.63\times10^{-3})^2\times2.14\times10^{-4}}{(3.68\times10^{-4})^2}$

$= 9.2 \times 10^{-2}$ mol dm^{-3} (1)

allow ecf on answer to part **(i)**

(iii) endothermic **because** K_c increases with temperature
mark is for explanation
allow ecf on answer to part **(ii)** (1) [5]

(d) (i) equilibrium moves to RHS (1)

more moles on RHS (1)

(ii) no change to equilibrium position (1)

[NOC*l*] and [NO] change by same amount (1) [4]

[Total: 15]

Methanol may be manufactured catalytically from *synthesis gas*, a mixture of CO, CO_2 and H_2. The CO is reacted with H_2 to form methanol, CH_3OH.

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g) \qquad \Delta H = -91 \text{ kJ mol}^{-1}$$

(c) From your understanding of Le Chatelier's principle, state **two** conditions that could be used in order to produce a high yield of methanol.

In **each** case, explain why the yield would increase.

condition 1 ....

explanation ....

....

condition 2 ....

explanation ....

.... [4]

Carbon monoxide, which can be used to make methanol, may be formed by reacting carbon dioxide with hydrogen.

$$CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g) \qquad K_c = 1.44 \text{ at } 1200\text{ K}$$

(d) **(i)** It has been suggested that, on a large scale, this reaction could be helpful to the environment.

Explain, with reasons, why this would be the case.

....

....

(ii) A mixture containing 0.50 mol of CO_2, 0.50 mol of H_2, 0.20 mol of CO and 0.20 mol of H_2O was placed in a 1.0 dm^3 flask and allowed to come to equilibrium at 1200 K.

Calculate the amount, in moles, of each substance present in the equilibrium mixture at 1200 K.

	CO_2	+	H_2	$\rightleftharpoons$	CO	+	H_2O
initial moles	0.50		0.50		0.20		0.20

(c) low temperature (1)
because forward reaction is exothermic (1)

high pressure (1)
because forward reaction goes to fewer molecules (1)
or shows a reduction in volume

increase [CO] **or** [H_2]
or remove CH_3OH (1)

correct explanation in terms of the effect of the change
on the position of equilibrium or on the rate of reaction (1)
(any two pairs)

(d) **(i)** removes CO_2 (1)
which causes greenhouse effect/global warming (1)

(ii)

	CO_2	+	H_2	$\rightleftharpoons$	CO	+	H_2O	
initial moles	0.50		0.50		0.20		0.20	
equil. moles	(0.50-x)		(0.50-x)		(0.20+x)		(0.20+x)	(1)
equil. concn.	$\frac{(0.50-x)}{1}$		$\frac{(0.50-x)}{1}$		$\frac{(0.20+x)}{1}$		$\frac{(0.20+x)}{1}$	

$$K_c = \frac{[CO][H_2O]}{[CO_2][H_2]}$$ (1)

$$K_c = \frac{(0.20+x)^2}{(0.50-x)^2} = 1.44$$ (1)

gives x = 0.18 (1)

at equilibrium,
$n(CO_2) = n(H_2) = 0.32$ **and**
$n(CO) = n(H_2O) = 0.38$ (1)

(b) Ethanoic acid, CH_3CO_2H, reacts with ethanol, C_2H_5OH, to produce ethyl ethanoate and water. The reaction is an example of dynamic equilibrium.

$$CH_3CO_2H + C_2H_5OH \rightleftharpoons CH_3CO_2C_2H_5 + H_2O$$

(i) Explain what is meant by *dynamic equilibrium*.

....

....

(ii) Write an expression for the equilibrium constant, K_c, for this reaction.

(c) A mixture of 6.0 g of ethanoic acid and 6.0 g of ethanol was added to 4.4 g of ethyl ethanoate and the overall mixture allowed to reach equilibrium. It was found that 0.040 mol of ethanoic acid was present in the equilibrium mixture.

(i) Calculate the number of moles of each compound, both initially and at equilibrium. Place the results in the spaces provided.

$$CH_3CO_2H \;+\; C_2H_5OH \;\rightleftharpoons\; CH_3CO_2C_2H_5 \;+\; H_2O$$

	CH_3CO_2H	C_2H_5OH	$CH_3CO_2C_2H_5$	H_2O
initially	....	....	....	0.00
at equilibrium	0.040	....	....	....

(ii) Calculate the equilibrium constant, K_c, for the reaction.

(iii) Explain why K_c in this reaction has no units.

....[4]

(b) (i) The forward reaction proceeds at the same rate as the reverse reaction (1)

(ii) $$K_c = \frac{[\text{ester}][H_2O]}{[\text{acid}][\text{alcohol}]}$$ (1)

(c) (i)

	acid	+ alcohol	⇌ ester	+ water
initially	$\frac{6}{60} = \underline{0.1}$	$\frac{6}{46} = \underline{0.13}$	$\frac{4.4}{88} = \underline{0.050}$	+ 0
				All three (1)
at equilm	$[0.040]$	$0.13 - 0.06 = \underline{0.07}$	$0.050 + 0.06 = \underline{0.11}$	$\underline{0.06}$
				all three (1)

(ii) $$K_c = \frac{0.11 \times 0.06}{0.04 \times 0.07} = 2.36$$ (1)

(iii) Units 'cancel' / two concn terms 'top + bottom' (1)

Total: 10

1 Hydrogen iodide dissociates into its elements according to the equation below.

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g)$$

(a) Write the expression for the equilibrium constant, K_c.

[1]

(b) At 120 °C the equilibrium mixture contains 1.47 mol dm^{-3} of HI(g), 0.274 mol dm^{-3} each of H_2(g) and I_2(g).

Calculate the value of K_c for the equilibrium at 120 °C.

[1]

(c) Suggest and explain why it would be more difficult to determine K_c for this equilibrium at room temperature.

....

....

....

....[2]

(d) (i) Explain how enthalpy changes, ΔH values, for covalent bonded molecules can be calculated from bond energies.

....

....

1 **(a)** $K_c = \dfrac{[H_2][I_2]}{[HI]^2}$ (1) **[1]**

(b) $K_c = \dfrac{0.274 \times 0.274}{(1.47)^2} = 0.035$ (1) **[1]**

(c) At room temperature:

iodine is a solid/solids not K_c expression (1)

$[I_2(g)]$ is small/concn too small to be measured (1)

it takes longer to reach equilibrium/reaction is slower (1) **[2 max]**

(d) **(i)** $\Delta H_{\text{reacn}} = \Delta H$ for bonds broken – ΔH for bonds made (1)

(ii) $2H - I \rightarrow H - H + I - I$

2 x 299 436 151 values (1)

$\Delta H = 2 \times 299 - (436 + 151)$

$= +11\ \text{kJ mol}^{-1}$ (1) **[3]**