

Tutorial letter 203/1/2016

General Chemistry IA (Theory)

CHE1501

Semester 1

Department of Chemistry

IMPORTANT INFORMATION:

This tutorial letter contains solutions to Assignment 03 .

BAR CODE

CONTENTS

1 KEY TO ASSIGNMENT 03.....3

Feedback on assignment 3:

- (1) a) A **hypothesis** is a *suggested* explanation for an observable phenomenon. It is often guesswork based on initial evidence, but has yet to be proven.

A **theory** is a tested, well-substantiated explanation for a phenomenon and is backed by evidence. A theory is valid as long as there is no evidence to dispute it. However, theories can be disproven if evidence is found to the contrary.

- b) A scientific **law** is a rigorous explanation of a phenomenon which is formed through repeated testing and evidence. At the time that the law is made, no exception has ever been found to that law.

In general, theories describe how matter behaves, but with laws one can usually explain why something behaves in a certain way.

- (2) a) $^{155}_{47}\text{Ag}$

- b) $^{86}_{37}\text{Rb}$

- c) $^{22}_{11}\text{Na}$

- (3) In general, main group elements have only one oxidation state, and these are written without Roman numerals, as their oxidation state is unambiguous.

The transition metals have more than one oxidation state, and therefore Roman numerals are required to clarify which oxidation state we're talking about. For example, iron may react with oxygen to form either FeO or Fe₂O₃. These are both called iron oxide, but in FeO, iron has an oxidation number of +2, while in Fe₂O₃, it has an oxidation number of +3. The rule for naming these compounds is to write the oxidation number of the metal after the name. Therefore, these compounds would be named iron (II) oxide and iron (III) oxide respectively.

However, there are exceptions to this guideline. Some main group metals can have more than one oxidation state include tin (Sn²⁺ or Sn⁴⁺) and lead (Pb²⁺ or Pb⁴⁺). Transition metals with only one common oxidation state include silver (Ag⁺), zinc (Zn²⁺), and cadmium (Cd²⁺). These don't require Roman numerals, but when in doubt, include the Roman numerals for all transition metals, as this would not be incorrect.

- (4) An orbit is a trajectory with well defined position and momentum at all points and time. In the Bohr model of an atom, an electron orbited the nucleus in a well defined circular path similarly to the way the earth orbits the sun.

However, an orbital in the quantum mechanical model is a three-dimensional region (a cloud) around the nucleus where the probability of finding an electron is at a maximum.

- (5) It is important to drink enough water, but drinking too much water can be harmful. The body's normal sodium level is 135-145mM (millimoles/liter). After drinking too much water, sodium levels become diluted, and below 120mM, dizziness and confusion set in. A concentration below 120mM can be fatal. This is why sports drinks contain sodium (electrolytes) to make sure that the sodium level is maintained / replaced when drinking a lot of water to prevent dehydration.
- (6) Read the section: "Effective Nuclear Charge" in the textbook; make sure you understand what it means and that you know the trends in the periodic table.
- (7) H_2 has two protons and two electrons involved in bonding. H_2^+ has two protons and one electron involved in bonding. The charge difference is greater in the bond with two electrons, making the bond of H_2 stronger than H_2^+ .
- (8) Read the "Chemistry and Life" box entitled "The Chemistry of Vision" in the textbook. (Page 410 in the 13th edition). This is more for general interest and won't be in the examination, as it contains line structure diagrams which you have not covered in CHE1501 and will only be done in CHE1502. However, it was included in the assignment as it is important to get an idea of how chemistry affects all aspects of everyday life.
- (9) Formal charge is given by the formula:

$$FC = V - N - (B/2) \text{ where:}$$

V = Number of valence electrons of the neutral atom in its ground state

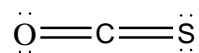
N = Number of non-bonding valence electrons on the atom

B = total number of electrons shared in bonds with other atoms

a) Number of valence electrons:

$$\begin{aligned} & \text{O} + \text{C} + \text{S} \\ & 6 + 4 + 6 = 16 \end{aligned}$$

Lewis Structure:



Oxidation numbers:

$$\begin{aligned} \text{O:} & \quad -2 \\ \text{C:} & \quad +4 \\ \text{S:} & \quad -2 \end{aligned}$$

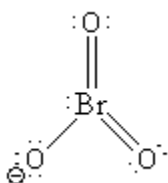
Formal Charges:

$$\begin{aligned} \text{O:} & \quad 6 - 4 - (4/2) = 0 \\ \text{C:} & \quad 4 - 0 - (8/2) = 0 \\ \text{S:} & \quad 6 - 4 - (4/2) = 0 \end{aligned}$$

b) Number of valence electrons:

$$\begin{aligned} & \text{Br} + \text{O} + \text{charge} \\ & 7 + (3 \times 6) + 1 = 26 \end{aligned}$$

Lewis Structure:



Oxidation numbers:

$$\begin{aligned} \text{O:} & \quad -2 \\ \text{Br:} & \quad +5 \end{aligned}$$

Formal Charges:

$$\begin{aligned} \text{O (double bonds):} & \quad 6 - 4 - (4/2) = 0 \\ \text{O (single bond):} & \quad 6 - 6 - (2/2) = -1 \\ \text{Br:} & \quad 7 - 2 - (10/2) = 0 \end{aligned}$$

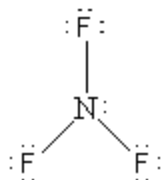
A very good YouTube video that explains why this Lewis structure is the correct one can be found at <http://www.youtube.com/watch?v=arkf8YDSgMM>

c) Number of valence electrons:

$$N + F$$

$$5 + (3 \times 7) = 26$$

Lewis Structure:



Oxidation numbers:

$$F: -1$$

$$N: +3$$

Formal Charges:

$$N: 5 - 2 - (6/2) = 0$$

$$F: 7 - 6 - (2/2) = 0$$

(10) a) i) $\cdot\ddot{\text{Cl}}\cdot$

Chlorine is in group 7 and therefore has 7 valence electrons in its ground state.

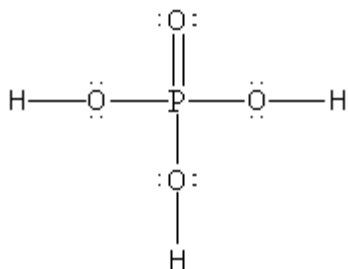
ii) $\cdot\ddot{\text{F}}\ominus$

Fluorine is in group 7 and therefore gains one electron to form a negatively charged fluorine ion with 8 valence electrons.

iii) $\cdot\ddot{\text{Br}}\text{---}\ddot{\text{Br}}\cdot$

Two bromine atoms combine to form a bromine molecule, and a pair of electrons is shared.

(10) b)



It is vital that you know how to draw Lewis diagrams. You should be able to show all lone pairs in these diagrams, assign formal charges to each atom, and indicate bond polarity on any polar bonds. The module CHE1502 assumes that you are 100% proficient in drawing Lewis diagrams, and you will not be able to pass that module without it. Now is an excellent time to make sure you know how, as these always feature in CHE1501 exams as well.

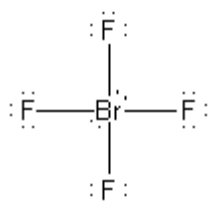
TRY THIS: For practice, try drawing the Lewis diagrams of some of the common acids - formic acid, acetic acid, phosphoric acid and sulphuric acid. Use lines for bonds and show all lone pairs. Try and determine the formal charge on each atom. Identify any polar bonds and indicate these on your diagrams.

(11) a) First, consider BrF_4^- :

As usual, we count the total number of valence electrons:

$$\begin{aligned} &\text{Br} + \text{F} + \text{charge} \\ &7 + (4 \times 7) + 1 = 36 \end{aligned}$$

We then, as always, place the least electronegative atom (bromine in this case) as the central atom, with the four fluorine atoms around it. Then we add the lone pairs on the fluorines and add up the total number of electrons: 8 from the single bonds, $4 \times 6 = 24$ from the lone pairs on the fluorines, giving a total of 32 electrons in total. But from our original calculation, there are 36 valence electrons in this molecule so we have to add another 4. Bromine is lower than group 2 in the periodic table and can therefore it is possible for it not to obey the octet rule. Therefore, we add the last four electrons to the Bromine giving the following Lewis structure:



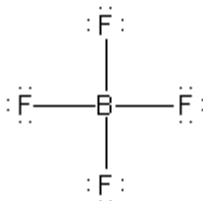
There are lone pairs on bromine, so BrF_4^- has an AX_4E_2 molecular shape, with 6 regions of electron density, four bonding pairs, and two lone pairs, making it **square planar** (see Table 3.2 on page 35 of the study guide)

Now consider BF_4^- :

Count the total number of valence electrons:

$$\begin{aligned} &\text{B} + \text{F} + \text{charge} \\ &3 + (4 \times 7) + 1 = 32 \end{aligned}$$

Place the least electronegative atom (boron in this case) as the central atom, with the four fluorine atoms around it. Then we add the lone pairs on the fluorines and add up the total number of electrons: 8 from the single bonds, $4 \times 6 = 24$ from the lone pairs on the fluorines, giving a total of 32 electrons in total. This adds up to the 32 valence electrons in our original calculation, so we are done, and the Lewis structure looks as follows:



There are no lone pairs on bromine, so BrF_4^- has an AX_4 molecular shape, with 4 regions of electron density, four bonding pairs, and zero lone pairs, making it **tetrahedral** (again, see Table 3.2 on page 35 of the study guide)

- b) Each of the molecules H_2O , H_2S , and H_2Se has two non-bonding lone pairs of electrons. As you move down the periodic table, the electron pairs are attracted by the nucleus less and less strongly. As a consequence of this lower attraction the lone pairs occupy more space, and this causes the bond angle of the H's to become more acute (smaller).

The bond angles are:

$$\begin{array}{ll} \text{H}_2\text{O:} & 104.5^\circ \\ \text{H}_2\text{S:} & 92.3^\circ \\ \text{H}_2\text{Se:} & 91.0^\circ \end{array}$$

TRY THIS: Can you tell the difference between ionic, polar covalent, non-polar covalent, coordinate-covalent and hydrogen bonds? What type of bond does KCl have? What about HCl? I_2 ? CuH?

(12) The density of a gas is given by the equation: $\rho = (PM)/(RT)$

Thus, we can see that the density of a gas is directly proportional to the molar mass of the gas; i.e. $\rho \propto M$.

The molar mass of water vapour (H_2O) is 18g/mol and for N_2 it is 28 g/mol. Therefore, water vapour is less dense than nitrogen gas.

(13) Rules to remember regarding oxidation states:

- **Elements in their elemental form have a 0 oxidation state.** ie. Br_2 , $Ag(s)$ etc.
- **Halogens are always -1, except when in their uncharged elemental form, or when bonded to another halogen.**
- **Oxygen is always -2, except in peroxides.** (Peroxides are of the form X_2O_2 , where X is a group 1 element. ie. H, Li, Na, K etc.)
- **Hydrogen is always +1, except when bonded to a metal.** ie. NaH

a) $Ni(OH)_2^-$

- Oxygen is always -2, except in peroxides, where it is -1.
- Hydrogen is always +1, except when bonded to a metal. In this case, it is bonded to oxygen, not to the nickel, so it is +1.
- From the oxidation states of O and H above, we see that each (OH) ligand has an overall charge of -1, and there are two (OH) ligands. Note that the $Ni(OH)_2^-$ molecule has an overall charge of -1. Therefore, Ni, must be +1, and **not** +2 as it would be in a neutral molecule.

O: -2
H: +1
Ni: +1

b) Na_2O_2

- Oxygen is always -2, except in peroxides, where it is -1.
- Hydrogen is always +1, except when bonded to a metal.
- The molecule is a peroxide.

Na: +1
O: -1

c) FeCl_6^{3-}

- Halogens are always -1, except when in their uncharged elemental form, or when bonded to another halogen.
- Each Cl ligand has an overall charge of -1, and there are six Cl ligands. Therefore, in order to make FeCl_6^{3-} have an overall charge of -3, Fe must be +3.

Cl: -1
Fe: +3

d) Br_2

- Elements in their elemental form have a 0 oxidation state. ie. Br_2 , Ag(s) etc.

Br: 0

e) BrF_4^+

- Hydrogen is always +1, except when bonded to a metal.
- **Fluorine can only have a -1 oxidation state.** If it is bonded to another halogen, the other halogen may take on a positive state, never the fluorine!
- Since there is a positive charge on the BrF_4^+ molecule, there must be a +5 charge on the bromine added to the four -1 charges of the fluorines to give the overall charge of +1.

F: -1
Br: +5

Additional Practice problems

Here are a few additional molecules which you may like to try on your own for practice:

Ag, HNO_3 , AgNO_3 , NO, H_2O , KMnO_4 , NH_3 , NO_2^- , $\text{Cu}(\text{CN})_2^-$, Na_2O_2 , NaH, K_2CrO_4 , Br_2



The first step when balancing redox reactions is **always** to assign the oxidation numbers first. If you cannot do this, or if you make a mistake with this step, you will not be able to identify either the oxidizing or reducing agents, nor will you be able to balance the equation using the half-reaction method.

Make sure that you understand Question 13 above thoroughly!

OK, here goes...

BrO₃:

O: -2

Br: +6

Do you see why?

3 oxygens give a total charge of $-2 \times 3 = -6$.

The molecule has an overall charge of 0.

Therefore, the bromium must be +6.

Br⁻: -1

Br₂: 0 (elemental state)

We can now see from the equation that bromium has been reduced from +6 to 0.

The Br⁻ ion has been oxidized from -1 to 0.

Remember that a reducing agent is itself oxidized!

Since Br⁻ was oxidized in the reaction, it is the reducing agent.

An oxidizing agent is itself reduced in a redox reaction!

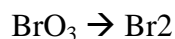
Since the bromium in BrO₃ was reduced in the reaction, BrO₃ is the oxidizing agent.

Write down the unbalanced half-reactions:

Oxidation (where oxidation takes place):

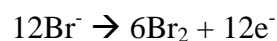


Reduction:



Balance these:

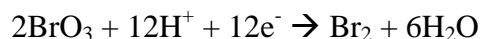
- Add water to balance the oxygen.
- Add H^+ ions to balance the hydrogen.
- Add electrons (e^-) to balance the charges.
- **NB! Remember to always add the electrons to the more positive side!**



Check: Left side has 12 bromines and -12 charge.

Right side has 12 bromines and -12 charge.

Thus, it is balanced.



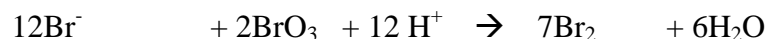
Check: Left side has 2 bromine, 6 oxygen, 12 hydrogen, and 0 charge.
(+12 from 12 H^+ , -12 from 12 e^-)

ie. +12 - 12 = 0

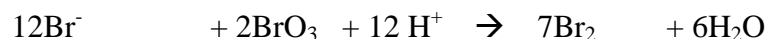
Right side has 2 bromine, 6 oxygen, 12 hydrogen, and 0 charge.

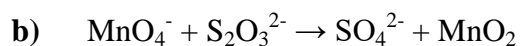
Thus, it is balanced.

Add them together:



The final equation in acid medium becomes:



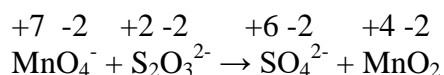


The steps in this type of problem are always the same:

Step 1:

Assign oxidation numbers.

Assign oxidation numbers to every atom in the equation:



Step 2:

Write the unbalanced oxidation and reduction half-reaction.

Identify what is being oxidized and what is being reduced, and write the unbalanced half-reactions:

The oxidation state of manganese changes from +7 to +4, so manganese is being reduced, and the unbalanced reduction half-reaction is:



The oxidation state of sulphur changes from +2 to +6, so sulphur is being oxidized, and the unbalanced oxidation half-reaction is:



The reducing *agent* is the substance which causes something else to be reduced, and is itself oxidized in a redox reaction. In this case, S in $\text{S}_2\text{O}_3^{2-}$ is being oxidized from a +2 to a +6 oxidation state, and causes the manganese in MnO_4^- to be reduced from a +7 to a +4 oxidation state. The reducing agent is therefore $\text{S}_2\text{O}_3^{2-}$.

The oxidizing *agent* is the substance which causes something else to be oxidized, and is itself reduced in a redox reaction. In this case, the Mn in MnO_4^- is being reduced from a +7 to a +4 oxidation state, and causes S in $\text{S}_2\text{O}_3^{2-}$ to be oxidized from a +2 to a +6 oxidation state. The oxidizing agent is therefore MnO_4^- .

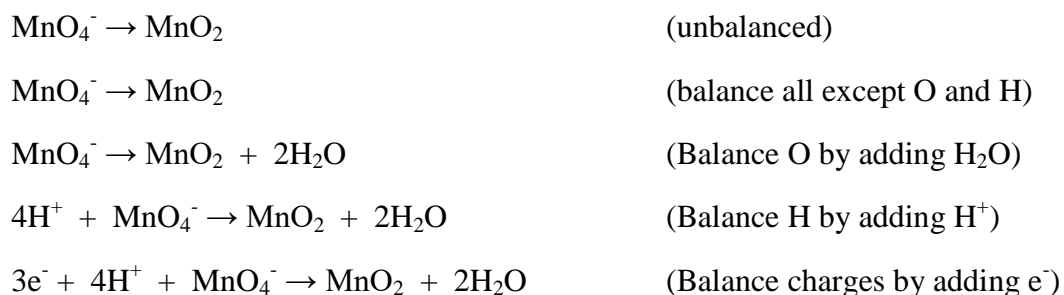
Step 3:**Balance the half reactions.**

The steps to balancing the half-reactions are as follows:

- First balance all atoms except oxygen and hydrogen by inspection.
- Secondly, balance oxygen *by adding water*.
- Thirdly, balance hydrogen by adding protons (H^+)
 - o Lastly, balance the charges by adding electrons (e^-) *to the more positive side*. (This is almost always the same side as the H^+ atoms.)
- Cancel out on both sides if necessary and write the balanced half-reaction.

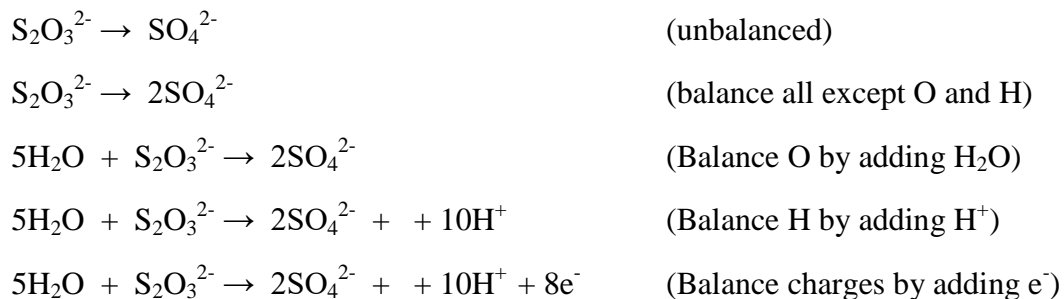
So we balance our half reactions by following the steps above:

Reduction half-reaction:



NB! Remember to always add the electrons to the more positive side!

Oxidation half-reaction:

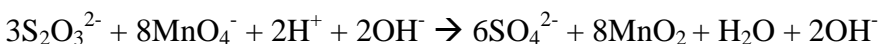


Check:

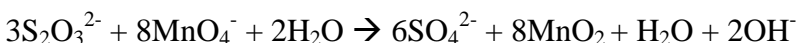
Left:		Right:	
S:	6	S:	6
O:	41	O:	41
H:	2	H:	2
Mn:	8	Mn:	8
Charge:	-12	Charge:	-12

Step 5: (*Only when balancing in basic medium*)**Add an equal amount of OH⁻ to both sides to cancel out H⁺.**

Always balance in acid medium first. If you wish to balance in basic medium, then add this step after you have balanced the reaction in acid medium.



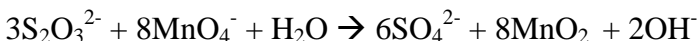
The H⁺ ions and OH⁻ ions combine to become water:



Cancel water molecules common to both sides:



The final balanced equation in basic medium is:

**Use the same steps in balancing all redox reactions!**

Check:

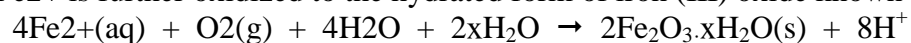
Left:		Right:	
S:	6	S:	6
O:	41	O:	41
H:	2	H:	2
Mn:	8	Mn:	8
Charge:	-12	Charge:	-12

15) It is very important to know how chemistry applies to everyday life. The chemistry behind the rusting of iron is as follows:

- First, iron is oxidized by oxygen to Fe^{2+} :



- The Fe^{2+} is further oxidized to the hydrated form of iron (III) oxide known as rust:



- Electrons cannot be transferred directly from iron to oxygen – they must travel through water.
- Ions in solution make a better conductor of electrons and carry electrical charge. Salt and acid dissociates in water to form ions.
- The more ions present, the faster rusting takes place.

16) Arrhenius acids are substances which increase $[\text{H}^{+}]$ in water.

Arrhenius bases are substances which increase $[\text{OH}^{-}]$ in water.

17) If HClO_4 is a strong acid, then its conjugate base, ClO_4^{-} must be a weak base.

- 19) I have noticed that people tend to omit this section. However, weak acid and base questions are critical to chemistry calculations in analytical chemistry, and you *must* know them. Just because they haven't appeared in the last few exams doesn't mean that they won't be asked in this paper. A weak acid or base question may well be present in the exam. For each question in Q19 I have put a detailed explanation of how to calculate the answer, and explained why we do each step. However, if a weak acid or base question is asked in the exam, you are not expected to write a detailed explanation like this. Therefore, I have also included a model answer for each question, which is similar to the answer that would be required from you in an exam.

a) **EXPLANATION:**

The first step is to figure out the balanced equation. Since we are dealing with a weak acid, we know that HOCl will lose a proton and form equilibrium with its conjugate base:



Next, we set up the ICE table (ICE stands for **I**nitial, **C**hange, **E**quilibrium), where all values are given in mol/L (M):

Concentration (M)	HOCl _(aq)	⇌	OCl ⁻ _(aq)	+	H ⁺ _(aq)
Initial					
Change					
Equilibrium					

The top line tells us what concentration we had before the equilibrium was reached. When we started, we had only 0.100 M of HOCl, and since dissociation has not yet taken place, there is not yet any OCl⁻ or H⁺. Therefore, our top line is simply our starting concentration:

Concentration (M)	HOCl _(aq)	⇌	OCl ⁻ _(aq)	+	H ⁺ _(aq)
Initial	0.100		-		-
Change					
Equilibrium					

The middle line represents the change that takes place when we allow the reaction to equilibrate (i.e. when dissociation takes place). Since HOCl is a weak acid, we don't know how much H⁺ dissociates from the HOCl. Let's call it x.

When x amount of H^+ dissociates from HOCl , the amount of HOCl decreases by x . The stoichiometry of this reaction (see the balanced equation) tells us that for each one mole of HOCl that dissociates, one mole of OCl^- and one mole of H^+ will be formed. Similarly, if HOCl decreases by $-x$ then OCl^- and H^+ will each increase by $+x$. We write this in the middle row of our ICE table:

Concentration (M)	$\text{HOCl}_{(\text{aq})}$	\rightleftharpoons	$\text{OCl}^-_{(\text{aq})}$	+	$\text{H}^+_{(\text{aq})}$
Initial	0.100		-		-
Change	$-x$		$+x$		$+x$
Equilibrium					

The third row is then simply the final amount of each species that exists at equilibrium:

Concentration (M)	$\text{HOCl}_{(\text{aq})}$	\rightleftharpoons	$\text{OCl}^-_{(\text{aq})}$	+	$\text{H}^+_{(\text{aq})}$
Initial	0.100		-		-
Change	$-x$		$+x$		$+x$
Equilibrium	$0.100 - x$		x		x

We know from our theory of chemical equilibrium that:

$$K_a = \frac{[\text{OCl}^-][\text{H}^+]}{[\text{HOCl}]}$$

By substituting the equilibrium concentration values from our ICE table:

$$K_a = \frac{(x)(x)}{(0.100 - x)} = \frac{x^2}{0.100 - x}$$

However, since HOCl is a weak acid, we know that only a small amount dissociates, and so the H^+ concentration will be much smaller than 0.100. Therefore, we can safely use the approximation that $0.100 - x \approx 0.100$.

This really simplifies our equation, as we can now write:

$$K_a = \frac{x^2}{0.100}$$

But K_a is given as 3.5×10^{-8} , so our equation becomes:

$$3.5 \times 10^{-8} = \frac{x^2}{0.100}$$

Therefore:

$$x^2 = (3.5 \times 10^{-8})(0.100)$$

$$x = 5.91 \times 10^{-5}$$

But we know from our ICE table that $x = [\text{H}^+]$. Therefore:

$$[\text{H}^+] = 5.91 \times 10^{-5} \text{ M}$$

$$\text{So: } \text{pH} = -\log(5.91 \times 10^{-5}) = 4.23$$

This looks about right, since as fairly dilute solution of a weak acid, we expect our answer to be a little below 7.

P.S. For those of you who are mathematicians and don't trust our assumption above, we can also solve the quadratic equation the long way, but arrive at the exact same answer. Let's try it:

$$K_a = 3.5 \times 10^{-8} = \frac{x^2}{0.100 - x}$$

$$x^2 + 3.5 \times 10^{-8} - 3.5 \times 10^{-9} = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-(3.5 \times 10^{-8}) \pm \sqrt{(3.5 \times 10^{-8})^2 - 4(1)(-3.5 \times 10^{-9})}}{2(1)} = 5.91 \times 10^{-5}$$

MODEL ANSWER

Concentration (M)	HOCl _(aq)	⇌	OCl ⁻ _(aq)	+	H ⁺ _(aq)
Initial	0.100		-		-
Change	-x		+x		+x
Equilibrium	0.100 - x		x		x

$$K_a = \frac{[\text{OCl}^{-}][\text{H}^{+}]}{[\text{HOCl}]} = \frac{x^2}{0.100 - x}$$

Assume x is small. Then $0.100 - x \approx 0.100$.

Then:

$$3.5 \times 10^{-8} = \frac{x^2}{0.100}$$

$$x^2 = (3.5 \times 10^{-8})(0.100)$$

$$x = 5.91 \times 10^{-5} \text{ M} = [\text{H}^{+}]$$

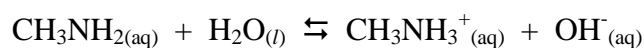
$$\text{pH} = -\log[\text{H}^{+}] = -\log(5.91 \times 10^{-5}) = 4.23$$

b) **EXPLANATION:**

In part b, we have a weak base. (The fact that we have a K_b value gives it away.)

I have noticed that where most people struggle with this equation is writing the balanced equation.

We know that, according to Bronsted-Lowry, a base will accept a proton. Arrhenius tells us that a base will increase the OH^- concentration in water. Therefore, if we have an aqueous solution of methylamine, we know that CH_3NH_2 will accept a proton from water and result in OH^- ions:



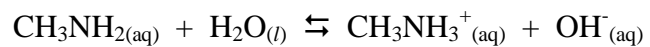
Recall from our theory of equilibrium that we do not include pure solvents (those with an (l) subscript showing the liquid phase)

Therefore, the equation can be written as:



You don't have to write it this way – just remember that water does not take part in the equilibrium because it is a pure solvent and therefore will also not appear in our ICE table.

The rest is similar to the procedure in (a), except that now we have OH^- in our answer, and not H^+ . We will therefore have to find the pH from the pOH.

MODEL ANSWER

Concentration (M)	$\text{CH}_3\text{NH}_{2(\text{aq})}$	\rightleftharpoons	$\text{CH}_3\text{NH}_3^+_{(\text{aq})}$	+	$\text{OH}^-_{(\text{aq})}$
Initial	1.0		-		-
Change	-x		+x		+x
Equilibrium	1.0 - x		x		x

$$K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} = \frac{x^2}{1.0 - x}$$

Assume x is small. Then $1.0 - x \approx 1.0$.

Then:

$$4.38 \times 10^{-4} = \frac{x^2}{1.0}$$

$$x^2 = (4.38 \times 10^{-4})(1.0)$$

$$x = 2.09 \times 10^{-2} \text{ M} = [\text{OH}^-]$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(2.09 \times 10^{-2}) = 1.68$$

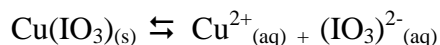
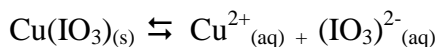
$$\text{pH} = 14 - 1.68 = 12.3$$

c) **EXPLANATION:**

When dealing with solubility, it's still the same principle, except that now it is the solid that doesn't appear in the ICE table.

Some people prefer to work with s (rather than x) for solubility, since it makes it easier to understand:

The question gives us a clue as to the oxidation number, since it says that it is copper(II) iodate, which means that copper has a 2+ charge. Since the molecule is neutral overall, this means that the iodate ion has a 2- charge:

**MODEL ANSWER**

Let s = solubility. Then:

Concentration (M)	Cu(IO₃)_(s)	⇌	Cu²⁺_(aq)	+	(IO₃)²⁻_(aq)
Initial	N/A		-		-
Change	N/A		+s		+s
Equilibrium	N/A		s		s

$$K_{sp} = [\text{Cu}^{2+}][(\text{IO}_3)^{2-}] = s^2$$

$$s^2 = 1.4 \times 10^{-7}$$

$$s = 3.7 \times 10^{-4} \text{ M}$$

TRY THIS:

I would strongly recommend that you try the weak acid and base problems from semester 2 (assignment 3) as well as these ones. They can be found on page 36 of tutorial letter 101, question 17(a) – 17(c).

Here are some additional practice exercises for you to try on topics not covered in this assignment:

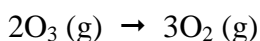
Q: Calculate the density of CO₂ gas at 0.940 atm and 38°C.

A: 1.77 g/L

Q: Is the rate of disappearance of reactants in a reaction always the same as the rate of appearance of products?

A: No. Why?

Q: Ozone, O₃, decomposes to molecular oxygen in the stratosphere according to the reaction:

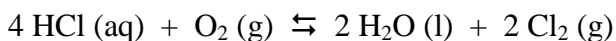


Would an increase in pressure favour the formation of ozone or of oxygen?

Would an increase in temperature favour the formation of ozone or of oxygen if this reaction is exothermic (ie. if it has a negative ΔH value)?

A: These are really easy – I'm sure I don't have to give you the answer to these! ☺

Q: Write the equilibrium constant expression for the reaction:



A: $K_c = \frac{[\text{Cl}_2]^2}{[\text{HCl}]^4[\text{O}_2]}$