Answers to Exercise 7.5
Equilibrium Constants and ICE Tables

1.

(a)	$AgBr_{(s)}$	$\rightleftharpoons$	$Ag^+_{(aq)}$	$+$	$Br^-_{(aq)}$
I(M)	n/a		0		0
C(M)	- x		+ x		+ x
E(M)	n/a		x		x

Step 1: Write a balanced chemical equation for the reaction

This was provided in the question. See above.

Step 2: Write the equilibrium constant expression for the reaction

$$K = \frac{\left(a_{Ag^+_{(aq)}}\right)\left(a_{Br^-_{(aq)}}\right)}{a_{AgBr_{(s)}}}$$

Step 3: Set up ICE table

See above.

This ICE table tracks the Initial, Change and Equilibrium molarities of the solutes. The Initial and Equilibrium amounts of AgBr do not matter (if there is enough available for the system to reach equilibrium) because it is a solid so its activity is always 1.

Step 4: Determine equilibrium activities for all reactants and products

$$a_{AgBr_{(s)}} = 1 \qquad\qquad a_{Ag^+_{(aq)}} = \frac{x\frac{mol}{L}}{1\frac{mol}{L}} = x$$

$$a_{Br^-_{(aq)}} = \frac{x\frac{mol}{L}}{1\frac{mol}{L}} = x$$

Step 5: Solve for x by substituting activities and equilibrium constant into equilibrium constant expression

$$5.0 \times 10^{-13} = \frac{(x)(x)}{1}$$

$$5.0 \times 10^{-13} = x^2$$

$$x = \sqrt{5.0 \times 10^{-13}}$$

$$x = 7.1 \times 10^{-7}$$

Step 6: Answer the question

$$M_{Ag^+_{(aq)}} = \left(a_{Ag^+_{(aq)}}\right)\left(1\frac{mol}{L}\right) = (7.1 \times 10^{-7})\left(1\frac{mol}{L}\right) = 7.1 \times 10^{-7}\frac{mol}{L}$$

$$M_{Br^-_{(aq)}} = \left(a_{Br^-_{(aq)}}\right)\left(1\frac{mol}{L}\right) = (7.1 \times 10^{-7})\left(1\frac{mol}{L}\right) = 7.1 \times 10^{-7}\frac{mol}{L}$$

Step 7: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

(b)   *This part of the question applies stoichiometry to the system. Calculate the mass of AgBr required to produce 1.00 L of $7.1 \times 10^{-7} \frac{mol}{L}$ $Ag^+_{(aq)}$ and $7.1 \times 10^{-7} \frac{mol}{L}$ $Br^-_{(aq)}$.*

Step 1: Calculate moles of Ag^+ (or of Br^-) produced

$n_{Ag^+} = \left(7.1 \times 10^{-7} \frac{mol}{L}\right)(1.00\ L) = 7.1 \times 10^{-7} mol$

Step 2: Use mole ratio to calculate moles of $AgBr$ required

$n_{AgBr} = 7.1 \times 10^{-7} mol\ Ag^+ \times \frac{1\ mol\ AgBr}{1\ mol\ Ag^+} = 7.1 \times 10^{-7} mol\ AgBr$

Step 3: Calculate mass of $AgBr$ required

$m_{AgBr} = 7.1 \times 10^{-7} mol \times \frac{187.772\ g}{1\ mol} = 1.3 \times 10^{-4} g$

$m_{AgBr} = 1.3 \times 10^{-4} g \times \frac{1000\ mg}{1\ g} = 0.13\ mg$

Step 4: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

Silver bromide has a low solubility in water (as evidenced by the small equilibrium constant for this reaction), so it makes sense that only a small mass is required to produce a saturated solution (i.e. to reach equilibrium).

2.	$CH_3CO_2H_{(aq)}$	$\rightleftharpoons$	$H^+_{(aq)}$	+	$CH_3CO^-_{2(aq)}$
I(M)	0.842		10^{-7}		0
C(M)	- x		+ x		+ x
E(M)	$0.842 - x \approx 0.842$		$x + 10^{-7} \approx x$		x

Step 1: Write a balanced chemical equation for the reaction

This was provided in the question. See above.

Step 2: Write the equilibrium constant expression for the reaction

$K = \frac{\left(a_{H^+_{(aq)}}\right)\left(a_{CH_3CO^-_{2(aq)}}\right)}{a_{CH_3CO_2H_{(aq)}}}$

Step 3: Calculate initial concentration of CH_3CO_2H

$n_{CH_3CO_2H} = 6.32\ g \times \frac{1\ mol}{60.0524\ g} = 0.105\ mol$

$M_{CH_3CO_2H} = \frac{0.105\ mol}{0.125\ L} = 0.842\ \frac{mol}{L}$

Step 4: Set up ICE table

This ICE table tracks the Initial, Change and Equilibrium molarities of the solutes.

In pure water at $25°C$, the concentration of H^+ is 10^{-7} M. Since the concentration of acetic acid is significantly higher than 10^{-7} M, it is reasonable to assume that this initial small amount of H^+ will be substantially smaller than the amount of H^+ generated by dissociation of the acid. Check this assumption at the end of the calculation!

The initial concentration of acetic is relatively high and it is a weak acid. It is therefore reasonable to assume that the majority of HA will remain in protonated form (in other words, that x will be substantially smaller than 0.25). Check this assumption at the end of the calculation!

Step 5: Determine equilibrium activities for all reactants and products

$$a_{CH_3CO_2H_{(aq)}} = \frac{(0.842-x)\frac{mol}{L}}{1\frac{mol}{L}} = 0.842 - x \qquad\qquad a_{H^+_{(aq)}} = \frac{x\frac{mol}{L}}{1\frac{mol}{L}} = x$$

$$a_{CH_3CO_2^-_{(aq)}} = \frac{x\frac{mol}{L}}{1\frac{mol}{L}} = x$$

Step 6: Solve for x by substituting activities and equilibrium constant into equilibrium constant expression

$$1.8 \times 10^{-5} = \frac{(x)(x)}{0.842}$$

$$(1.8 \times 10^{-5})(0.842) = x^2$$

$$x = \sqrt{(1.8 \times 10^{-5})(0.842)}$$

$$x = 0.0039$$

Step 7: Check assumptions. If invalid, repeat calculation without the invalid assumption.

#1: $10^{-7} + x \approx x$ $10^{-7} + 0.0039 = 0.0039$ assumption is okay

#2: $0.842 - x \approx 0.842$ $0.842 - 0.0039 = 0.8381 = 0.84$ assumption is okay

(You get exactly the same answer for x if you don't make assumption #2.)

Step 8: Answer the question

$$M_{H^+_{(aq)}} = \left(a_{H^+_{(aq)}}\right)\left(1\frac{mol}{L}\right) = (3.9 \times 10^{-3})\left(1\frac{mol}{L}\right) = 3.9 \times 10^{-3}\frac{mol}{L}$$

$$M_{CH_3CO_2^-_{(aq)}} = \left(a_{CH_3CO_2^-_{(aq)}}\right)\left(1\frac{mol}{L}\right) = (3.9 \times 10^{-3})\left(1\frac{mol}{L}\right) = 3.9 \times 10^{-3}\frac{mol}{L}$$

$$M_{CH_3CO_2H_{(aq)}} = \left(a_{CH_3CO_2H_{(aq)}}\right)\left(1\frac{mol}{L}\right) = (0.842 - 0.0039)\left(1\frac{mol}{L}\right) = 0.838\frac{mol}{L}$$

Step 9: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

This was a relatively concentrated solution of a weak acid. A small amount of it dissociated but most of the acid remained protonated. There was still enough dissociation to lower the pH of the solution to 2.4.

3.	$PCl_{5(g)}$	$\rightleftharpoons$	$PCl_{3(g)}$	$+$	$Cl_{2(g)}$
I(bar)	1.97		0		0
C(bar)	- x		+ x		+ x
E(bar)	$1.97 - x$		x		x

Step 1: Write a balanced chemical equation for the reaction

This was provided in the question. See above.

Step 2: Write the equilibrium constant expression for the reaction

$$K = \frac{\left(a_{PCl_{3(g)}}\right)\left(a_{Cl_{2(g)}}\right)}{a_{PCl_{5(g)}}}$$

Step 3: Convert temperature to Kelvin

$T = 200\,°C = 473.15\,K$ 　　　　　　*pressure of gases depend on temperature...*

Step 4: Calculate initial partial pressure of PCl_5

$PV = nRT$

$$P = \frac{nRT}{V} = \frac{(0.125\ mol)\left(8.314\,462\,\frac{Pa·m^3}{mol·K}\right)(473.15\ K)}{2.50\ L} \times \frac{1000\ L}{1\ m^3} \times \frac{1\ bar}{10^5\ Pa} = 1.97\ bar$$

Step 5: Set up ICE table

See above.

This ICE table tracks the Initial, Change and Equilibrium partial pressures of the gases. Partial pressures are used because activities of gases are calculated by dividing the pressure by 1 bar.

Step 6: Determine equilibrium activities for all reactants and products

$a_{PCl_{5(g)}} = 1.97 - x$ 　　　　　　　　　　　　$a_{PCl_{3(g)}} = \frac{x\ bar}{1\ bar} = x$

　　　　　　　　　　　　　　　　　　　　　　$a_{Cl_{2(g)}} = \frac{x\ bar}{1\ bar} = x$

Step 7: Solve for x by substituting activities and equilibrium constant into equilibrium constant expression

$0.155 = \frac{(x)(x)}{1.97-x}$

$(0.155)(1.97 - x) = x^2$

$x^2 + 0.155x - 0.305 = 0$

Solve for x using the quadratic equation or an equation solver (if your calculator has that feature).

$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ 　　　　*where a = 1, b = 0.155 and c = −0.305*

$x = \frac{-0.155 \pm \sqrt{(0.155)^2 - 4(1)(-0.305)}}{2(1)}$

$x = 0.480\ or\ x = -0.635$

It's not possible to have a negative activity (and the activities of the two products are both equal to x) therefore $x = 0.480$.

Step 8: Answer the question

$$p_{PCl_{3(g)}} = \left(a_{PCl_{3(g)}}\right)(1 \; bar) = (0.480)(1 \; bar) = 0.480 \; bar$$

$$p_{Cl_{2(g)}} = \left(a_{Cl_{2(g)}}\right)(1 \; bar) = (0.480)(1 \; bar) = 0.480 \; bar$$

$$p_{PCl_{5(g)}} = \left(a_{PCl_{5(g)}}\right)(1 \; bar) = (1.97 - 0.480)(1 \; bar) = 1.49 \; bar$$

Step 9: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

The equilibrium constant was relatively close to 1 (within an order of magnitude) therefore the pressures of products and reactants have similar orders of magnitude.

The equilibrium constant was smaller than 1 so, at equilibrium, the activities of products are lower than the activities of reactants.

4.

	$H_{2(g)}$	+	$I_{2(g)}$	$\rightleftharpoons$	$2 \; HI_{(g)}$
I(bar)	2.2		2.2		0
C(bar)	- x		- x		+ 2 x
E(bar)	$2.2 - x$		$2.2 - x$		2 x

Step 1: Write a balanced chemical equation for the reaction

This was provided in the question. See above.

Step 2: Write the equilibrium constant expression for the reaction

$$K = \frac{\left(a_{HI_{(g)}}\right)^2}{\left(a_{H_{2(g)}}\right)\left(a_{I_{2(g)}}\right)}$$

Step 3: Convert temperature to Kelvin

$T = 250 \; °C = 523.15 \; K$ *pressures of gases depend on temperature...*

Step 4: Calculate initial partial pressures of H_2 and I_2

$PV = nRT$

$$P = \frac{nRT}{V} = \frac{(0.25 \; mol)\left(8.314\,462 \frac{Pa \cdot m^3}{mol \cdot K}\right)(523.15 \; K)}{5.00 \; L} \times \frac{1000 \; L}{1 \; m^3} \times \frac{1 \; bar}{10^5 \; Pa} = 2.2 \; bar$$

Since there were equal moles of each gas, they have equal partial pressures.

Step 5: Set up ICE table

See above.

This ICE table tracks the Initial, Change and Equilibrium partial pressures of the gases. Partial pressures are used because activities of gases are calculated by dividing the pressure by 1 bar.

Step 6: Determine equilibrium activities for all reactants and products

$$a_{H_{2(g)}} = \frac{(2.2-x)\ bar}{1\ bar} = 2.2 - x \qquad\qquad a_{HI_{(g)}} = \frac{2x\ bar}{1\ bar} = 2x$$

$$a_{I_{2(g)}} = \frac{(2.2-x)\ bar}{1\ bar} = 2.2 - x$$

Step 7: Solve for x by substituting activities and equilibrium constant into equilibrium constant expression

$$122 = \frac{(2x)^2}{(2.2-x)(2.2-x)}$$

$$122 = \frac{(2x)^2}{(2.2-x)^2}$$

$$\sqrt{122} = \frac{2x}{2.2-x}$$

$$\sqrt{122}(2.2 - x) = 2x$$

$$2.2\sqrt{122} = 2x + \sqrt{122}\,x$$

$$x = \frac{2.2\sqrt{122}}{2+\sqrt{122}} = 1.8$$

Step 8: Answer the question

$$p_{HI_{(g)}} = \left(a_{HI_{(g)}}\right)(1\ bar) = (2 \times 1.8)(1\ bar) = 3.7\ bar$$

$$p_{H_{2(g)}} = \left(a_{H_{2(g)}}\right)(1\ bar) = (2.2 - 1.8)(1\ bar) = 0.3\ bar$$

$$p_{I_{2(g)}} = \left(a_{I_{2(g)}}\right)(1\ bar) = (2.2 - 1.8)(1\ bar) = 0.3\ bar$$

Step 9: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

The equilibrium constant was relatively close to 1 (within two orders of magnitude) therefore the pressures of products and reactants have similar orders of magnitude.

The equilibrium constant was larger than 1 so, at equilibrium, the activities of products are higher than the activities of reactants.