A saturated solution of iodine in water contains 0.330 g $I_2$ per litre, but more than this amount can dissolve in a potassium iodide solution because of the following equilibrium.

$$I^-(aq) + I_2(aq) \rightleftharpoons I_3^-(aq)$$

A 0.100 M KI solution dissolves 12.5 g of $I_2$ per litre, most of which is converted to $I_3^-$ (aq). Assuming that the concentration of $I_2(aq)$ in all saturated solutions is the same, calculate the equilibrium constant for the above reaction.

The molar mass of $I_2$ is $(2 \times 126.90) \text{ g mol}^{-1} = 253.8 \text{ g mol}^{-1}$.

As 0.330 g of $I_2$ dissolves in a litre of water, the concentration of $I_2$ in the saturated solution of iodine in water is therefore:

$$[I_2(aq)] = \frac{0.330}{253.8} = 0.0013 \text{ M}$$

12.5 g of $I_2$ corresponds to $\frac{12.5g}{253.8\text{ g mol}^{-1}} = 0.0493\text{ mol}$ and when this dissolves in a litre of KI solution, the concentration is 0.0493 M.

For the equilibrium, the reaction table is therefore:

<table>
<thead>
<tr>
<th></th>
<th>$I^-(aq)$</th>
<th>$I_2(aq)$</th>
<th>$I_3^-(aq)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>initial</td>
<td>0.100</td>
<td>0.0493</td>
<td>0</td>
</tr>
<tr>
<td>change</td>
<td>$-x$</td>
<td>$-x$</td>
<td>$+x$</td>
</tr>
<tr>
<td>equilibrium</td>
<td>0.100-$x$</td>
<td>0.0493-$x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

Assuming that $[I_2(aq)]$ is the same as in the saturated solution (as stated in the question), $0.0493 - x = 0.0013$ so $x = 0.048$ giving:

$$[I^-(aq)] = 0.100 - 0.048 = 0.052 \text{ M}, [I_2(aq)] = 0.0013 \text{ M} \text{ and } [I_3^-(aq)] = 0.048 \text{ M}.$$  

The equilibrium constant is therefore:

$$K_c = \frac{[I_3^-(aq)]}{[I^-(aq)][I_2(aq)]} = \frac{(0.048)}{(0.052)(0.0013)} = 710$$

Answer: 710
• When 1.0 mol of acetic acid and 1.0 mol of ethanol are mixed they react according to the following equation.

\[
C_2H_5OH(l) + CH_3COOH(l) \rightleftharpoons CH_3COOC_2H_5(l) + H_2O(l)
\]

At equilibrium the mixture contains 0.67 mol of the ester (CH₃COOC₂H₅). What is the equilibrium constant for the reaction?

At equilibrium, there are \((1.0 - 0.67)\) mol of C₂H₅OH and CH₃COOH together with 0.67 mol of CH₃COOC₂H₅ and of H₂O(l). The equilibrium constant is therefore:

\[
K_c = \frac{[CH_3COOC_2H_5(l)][H_2O(l)]}{[C_2H_5OH(l)][CH_3COOH(l)]} = \frac{(0.67)(0.67)}{(1.0 - 0.67)(1.0 - 0.67)} = 4.1
\]

Note that although amounts (mol) rather than concentrations are known, the volume is the same for each species and cancels through in the equation. The reaction involves mixtures of liquids and not aqueous solution so the concentration of water is included in the expression.

ANSWER: 4.1

• Bromine-containing compounds are even more ozone depleting than the analogous chlorine-containing compounds. In the stratosphere, an equilibrium exists between bromine and the NOₓ species. One of these equilibrium reactions is:

\[
2NOBr(g) \rightleftharpoons 2NO(g) + Br_2(g)
\]

To study this reaction, an atmospheric chemist places a known amount of NOBr in a sealed container at 25 °C to a pressure of 0.250 atm and observes that 34% of it decomposes into NO and Br₂. What is \(K_p\) for this reaction?

The initial partial pressure of NOBr is 0.250 atm. At equilibrium, 34% of it has decomposed so its partial pressure at equilibrium is 0.66 × 0.25 = 0.165 atm.

The reaction stoichiometry means that 2 moles of NO(g) are produced for every 2 moles of NOBr(g) that decompose. The equilibrium partial pressure of NO is therefore \(0.34 \times 0.25 = 0.085\) atm.

1 mole of Br₂(g) is produced for every 2 moles of NOBr(g) that decompose. The equilibrium partial pressure of Br₂(g) is therefore \(\frac{1}{2} \times 0.34 \times 0.25 = 0.0425\) atm.

The equilibrium constant in terms of partial pressures, \(K_p\), is therefore:

\[
K_p = \frac{(P_{NO})^2(P_{Br_2})}{(P_{NOBr})^2} = \frac{(0.085)^2(0.045)}{(0.165)^2} = 0.0113
\]

ANSWER: 0.0113
The value of the equilibrium constant, $K_c$, for the following reaction is 0.118 mol L$^{-1}$.

$$2\text{CO}_2(\text{g}) + \text{N}_2(\text{g}) \rightleftharpoons 2\text{CO}(\text{g}) + 2\text{NO}(\text{g})$$

What is the equilibrium concentration of CO(g) if the equilibrium concentration of $[\text{CO}_2(\text{g})] = 0.392 \text{ M}$, $[\text{N}_2(\text{g})] = 0.419 \text{ M}$ and $[\text{NO}(\text{g})] = 0.246 \text{ M}$?

The equilibrium constant in terms of concentrations, $K_c$, is:

$$K_c = \frac{[\text{CO}(\text{g})]^2[\text{NO}(\text{g})]^2}{[\text{CO}_2(\text{g})]^2[\text{N}_2(\text{g})]} = \frac{[\text{CO}(\text{g})]^2(0.246)^2}{(0.392)^2(0.419)} = 0.118$$

Therefore, $[\text{CO}(\text{g})]^2 = 0.126 \text{ M}^2$ or $[\text{CO}(\text{g})] = 0.354 \text{ M}$

Answer: 0.354 M

When hydrogen cyanide (HCN) is dissolved in water it dissociates into ions according to the following equation.

$$\text{HCN}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{CN}^-(\text{aq})$$

The equilibrium constant for this reaction is $K_c = 6.2 \times 10^{-10} \text{ mol L}^{-1}$. If 1.00 mol of HCN is dissolved to make 1.00 L of solution, calculate the percentage of HCN that will be dissociated.

The initial concentration of HCN is 1.00 M. The reaction table is:

<table>
<thead>
<tr>
<th></th>
<th>$\text{HCN}(\text{aq})$</th>
<th>$\text{H}^+(\text{aq})$</th>
<th>$\text{CN}^-(\text{aq})$</th>
</tr>
</thead>
<tbody>
<tr>
<td>[initial]</td>
<td>1.00</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>change</td>
<td>$-x$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td>[equilibrium]</td>
<td>1.00-$x$</td>
<td>$x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

$$K_c = \frac{[\text{H}^+(\text{aq})][\text{CN}^-(\text{aq})]}{[\text{HCN}(\text{aq})]} = \frac{(x)(x)}{(1.00-x)(1.00-x)} = \frac{x^2}{(1.00-x)} = 6.2 \times 10^{-10}$$

As $K_c$ is very small, the amount of dissociation will be tiny and $1.00-x \sim 1.00$. Hence,

$$x^2 \sim 6.2 \times 10^{-10}x, \text{ so } x = 2.5 \times 10^{-5} = [\text{H}^+(\text{aq})] = [\text{CN}^-(\text{aq})].$$

As $[\text{HCN}(\text{aq})] = 1.00 - x$, $[\text{HCN}(\text{aq})] = 1.00$ and the percentage dissociation is:

$$\text{percentage dissociation} = \frac{2.5 \times 10^{-5}}{1.00} \times 100 = 2.5 \times 10^{-3} \%$$

Answer: 2.5 $\times 10^{-3}$ %