1. (a) K₂[PtF₆]

Potassium hexafluoridoplatinate(IV)

(b) [CoCl₂(NH₃)₄]Cl·2H₂O

Tetramminedichloridocobalt(III) chloride – 2 – water.

2. (a) tetraamminezinc(II) sulfate-2-water

[Zn(NH₃)₄]SO₄·2H₂O

(b) tetraaquaoxalatochromium(III) ion

[Cr(C₂O₄)(H₂O)₄]⁺

3. [Pt(NH₃)₂Cl₂] can exist as two geometric isomers:

\[
\begin{align*}
&\text{cis (cisplatin)} & \text{trans (transplatin)} \\
&\quad \begin{array}{c}
\text{H₃N} \\
\text{Pt} \\
\text{Cl} \\
\text{H₃N} \\
\text{Cl}
\end{array} & \begin{array}{c}
\text{Cl} \\
\text{Pt} \\
\text{NH₃} \\
\text{Cl} \\
\text{H₃N}
\end{array}
\end{align*}
\]

4. Coordination isomers have a ligand and a counter ion exchanged.

[Cr(H₂O)₅Cl]SO₄ has a Cl⁻ ion as a ligand and SO₄²⁻ as a counter ion.

(c) [Cr(H₂O)₅SO₄]Cl is a coordination isomer: it has a SO₄²⁻ ion as a ligand and Cl⁻ as a counter ion. (Note that SO₄²⁻ coordinates to a metal using oxygen atoms rather than the sulfur.)

5. (a) Mn²⁺ [Ar] (4s)²(3d)⁵

(b) Cr [Ar] (4s)¹(3d)⁵

(c) Ni²⁺ [Ar] (3d)⁸

(d) Fe [Ar] (4s)²(3d)⁶

(e) Fe³⁺ [Ar] (3d)⁵

(f) Cu²⁺ [Ar] (3d)⁹

(g) Zn²⁺ [Ar] (3d)¹⁰

6. (a) Between experiments (2) and (4), [Fe²⁺] and [O₂] are unchanged. [H⁺] is doubled and this leads the rate to double. The reaction is first-order with respect to [H⁺].

Between experiments (3) and (4), [Fe²⁺] and [H⁺] are unchanged. [O₂] is doubled and this doubles the rate: the reaction is first-order with respect to [O₂].

Between experiments (1) and (3), [O₂] is unchanged. [Fe²⁺] and [H⁺] are both doubled and this leads to the rate increasing by a factor of 16. As the
reaction is first-order with respect to [H\(^+\)], the rate would double because of the doubling in [H\(^+\)]. The doubling in [Fe\(^{2+}\)] therefore increases the rate by a factor of 8: the reaction is third-order (\(2^3 = 8\)) with respect to [Fe\(^{2+}\)]. Overall,

\[
\text{rate} = k[\text{Fe}^{2+}]^3[\text{O}_2][\text{H}^+]
\]

(b) Using the first experiment, [Fe\(^{2+}\)] = [O\(_2\)] = 1 \times 10^{-3} \text{ M} and [H\(^+\)] = 0.1 \text{ M}. The rate is 5 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}. Using the rate equation from (a):

\[
\text{rate} = k(1 \times 10^{-3} \text{ M})^3 \times (1 \times 10^{-3} \text{ M}) \times (0.1 \text{ M}) = 5 \times 10^{-4} \text{ M s}^{-1}
\]

Hence, \(k = 5 \times 10^9 \text{ M}^{-4} \text{ s}^{-1}\). The same value is obtained using the data from the other experiments.

The units of \(k\) are obtained by matching the units on the right and left-hand side of the rate equation.

The rate is the change in concentration with time and has units of “M s\(^{-1}\”).

On the left-hand side of the rate law, [Fe\(^{2+}\)]^3[O\(_2\)][H\(^+\)] has units of \(\text{M}^5\).

To match, \(k\) must have units of “M\(^{-4}\) s\(^{-1}\)”:

\[
\text{M s}^{-1} = (\text{M}^4 \text{ s}^{-1}) \times \text{M}^5
\]

(c) From the chemical equation, 4 mol of Fe\(^{3+}\) are made for every 1 mol of O\(_2\) that is lost: rate of formation of Fe\(^{3+}\) = 4 \times rate of loss of O\(_2\).

In experiment (3), the rate of loss of O\(_2\) is 8 \times 10^{-3} \text{ M s}^{-1} so the rate of formation of Fe\(^{3+}\) = 4 \times (8 \times 10^{-3}) = 3 \times 10^{-2} \text{ M s}^{-1}.

(d) When [Fe\(^{2+}\)] = [O\(_2\)] = 4 \times 10^{-3} \text{ M} and [H\(^+\)] = 0.1 \text{ M}, the rate law gives:

\[
\text{rate} = k(4 \times 10^{-3} \text{ M})^3 \times (4 \times 10^{-3} \text{ M}) \times (0.1 \text{ M}) = 0.128 \text{ M s}^{-1}
\]

As 4 mol of Fe\(^{2+}\) are consumed for every 1 mol of O\(_2\), the rate of loss of Fe\(^{2+}\) is (4 \times 0.128) \text{ M s}^{-1} = 0.5 \text{ M s}^{-1}.

7. (a) As \(t_{1/2} = \frac{\ln 2}{k}\), \(k = \frac{\ln 2}{t_{1/2}} = \frac{\ln 2}{6.00 \times 10^4 \text{ s}^{-1}} = 1.16 \times 10^{-5} \text{ s}^{-1}\)

(b) For a first-order reaction,
\[
\ln \frac{[A]}{[A_0]} = -kt
\]

After 1 hour = \((60 \times 60)\) s = 3600 s, the fraction that remain will be:

\[
\frac{[A]}{[A_0]} = e^{-(1.16 \times 10^{-5} \times 3600)} = 0.959
\]

The percentage that will have reacted is 4.1%.

8. The Arrhenius equation relates the rate constant with the temperature:

\[
k = A \ e^{\frac{E_a}{RT}}
\]

If the rate constant is known at two temperatures, this can be rewritten as:

\[
\ln \left( \frac{k_2}{k_1} \right) = \frac{E_a}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)
\]

At 30 °C, \(T = (273 + 30) = 303\) K and \(k \) is \(5.1 \times 10^6\) s\(^{-1}\). At 50 °C, \(T = (273 + 50) = 323\) K and \(k \) is \(1.9 \times 10^7\) s\(^{-1}\):

\[
\ln \left( \frac{1.9 \times 10^7 \text{ s}^{-1}}{5.1 \times 10^6 \text{ s}^{-1}} \right) = \frac{E_a}{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})} \left( \frac{1}{(303 \text{ K})} - \frac{1}{(323 \text{ K})} \right)
\]

Hence \(E_a = 54000\) J mol\(^{-1}\) = 54 kJ mol\(^{-1}\)

This is a fairly typical activation energy and presumably corresponds, in the present case, to the partial breaking of the weak N-N bond in \(\text{N}_2\text{O}_4\). An activation energy of this approximate size is quite common and leads to the ‘rule of thumb’ that the rate is doubled by increasing the temperature by 10 °C. In this example, the temperature increases from 30 °C to 50 °C and the rate quadruples.

At either temperature, \(k = A \ e^{\frac{E_a}{RT}}\) so using \(k = 5.1 \times 10^6\) s\(^{-1}\) at \(T = 303\) K:

\[5.1 \times 10^6 \text{ s}^{-1} = A \times e^{(54000)/(8.314 \times 303)}\] so \(A = 8.6 \times 10^{15}\) s\(^{-1}\).