c) What is the oxidation number of Cr in Cr^{3+} ?

This one is obvious: the oxidation number (charge on the atom) is +3.

Conclusion: The oxidation number of a monatomic ion is the charge on the ion.

d) What is the oxidation number of S in SO_4^{2-} ?

	S	O_4	$^{2-}$
individual charge for an atom ⟶	x	−2	
total charge (all atoms) ⟶	x	−8	= −2

The ion has a 2− charge overall, requiring the solution of the equation:

$x - 8 = -2$, which gives $x = +6$.

Therefore, **the oxidation number of S is +6.**

***EXERCISES*:**

3. Calculate the oxidation number of the atom in **bold** type.

a) $H\mathbf{N}O_3$	e) $\mathbf{N}H_4^+$	i) $\mathbf{Al}(OH)_4^-$	m) $H\mathbf{Cl}O_3$	q) $K_2\mathbf{U}O_4$
b) $\mathbf{N}O_2^-$	f) $\mathbf{N}_3^-$	j) $\mathbf{S}_2F_{10}$	n) $\mathbf{N}_2H_5^+$	r) $\mathbf{C}_3H_6O$
c) $\mathbf{Cr}O_4^{2-}$	g) $\mathbf{C}_2H_6$	k) $\mathbf{N}_2O_3$	o) $\mathbf{N}H_2OH$	s) $\mathbf{S}_8$
d) $\mathbf{Cr}_2O_7^{2-}$	h) $\mathbf{C}_3H_8$	l) $H\mathbf{Cl}O_4$	p) $\mathbf{C}_2O_4^{2-}$	t) $\mathbf{C}_4H_6$

4. Assign oxidation numbers to the **bold** species in each of the following unbalanced reaction equations. Then determine which species undergoes oxidation in each reaction.

a) $\mathbf{Cl}O_2 + \mathbf{C} \longrightarrow \mathbf{Cl}O_2^- + \mathbf{C}O_3^{2-}$

b) $\mathbf{Sn}^{2+} + Cl^- + \mathbf{Br}O_3^- \longrightarrow \mathbf{Sn}Cl_6^{2-} + \mathbf{Br}^-$

c) $\mathbf{Mn}O_4^- + \mathbf{C}_2O_4^{2-} \longrightarrow \mathbf{Mn}O_2 + \mathbf{C}O_2$

d) $NO_3^- + H_2\mathbf{Te} \longrightarrow NO + \mathbf{Te}O_4^{2-}$

5. Which of the following are redox reactions?

a) $I_2 + 5\,HOBr + H_2O \longrightarrow 2\,IO_3^- + 5\,Br^- + 7\,H^+$

b) $4\,Ag^+ + Cr_2O_7^{2-} + H_2O \longrightarrow 2\,Ag_2CrO_4 + 2\,H^+$

c) $KHCO_3 + HI \longrightarrow KI + CO_2 + H_2O$

d) $2\,H_2O \longrightarrow 2\,H_2 + O_2$

e) $H_2SO_4 + BaCl_2 \longrightarrow BaSO_4 + 2\,HCl$

f) $Fe + H_2SO_4 \longrightarrow FeSO_4 + H_2$

SNEAKY TRICK!

You will sometimes find that the solution to a problem only requires you to determine which species have been oxidized and which have been reduced. Look at the following oxidations and reductions.

i) $ClO_3^- \longrightarrow ClO_4^-$ (oxidation: $Cl^{5+} \longrightarrow Cl^{7+}$)

ii) $H_2O_2 \longrightarrow H_2O$ (reduction: $O^- \longrightarrow O^{2-}$)

iii) $Cr^{3+} \longrightarrow CrO_4^{2-}$ (oxidation: $Cr^{3+} \longrightarrow Cr^{6+}$)

iv) $NO_2 \longrightarrow N_2O_3$ (reduction: $N^{4+} \longrightarrow N^{3+}$)

In each of these cases, **THE NUMBER OF ATTACHED OXYGEN ATOMS INCREASES DURING AN OXIDATION,** and **THE NUMBER OF ATTACHED OXYGENS DECREASES DURING A REDUCTION.** In the last example, (iv), the number of oxygens go from 2 O's per N–atom to 1.5 O's per N–atom.

ANSWERS TO UNIT V : ELECTROCHEMISTRY

1. (a) $Mn \longrightarrow Mn^{2+} + 2\,e^-$; Mn is oxidized and is the reducing agent.
   $Hg^{2+} + 2\,e^- \longrightarrow Hg$; Hg^{2+} is reduced and is the oxidizing agent.

   (b) $H_2 \longrightarrow 2\,H^+ + 2\,e^-$; H_2 is oxidized and is the reducing agent.
   $Sn^{4+} + 2\,e^- \longrightarrow Sn^{2+}$; Sn^{4+} is reduced and is the oxidizing agent.

   (c) $Li \longrightarrow Li^+ + e^-$; Li is oxidized and is the reducing agent.
   $F_2 + 2\,e^- \longrightarrow 2\,F^-$; F_2 is reduced and is the oxidizing agent.

   (d) $Cr^{2+} \longrightarrow Cr^{3+} + e^-$; Cr^{2+} is oxidized and is the reducing agent.
   $Br_2 + 2\,e^- \longrightarrow 2\,Br^-$; Br_2 is reduced and is the oxidizing agent.

   (e) $Fe^{2+} \longrightarrow Fe^{3+} + e^-$; Fe^{2+} is oxidized and is the reducing agent.
   $Sn^{4+} + 2\,e^- \longrightarrow Sn^{2+}$; Sn^{4+} is reduced and is the oxidizing agent.

2. (a) $2\,Cs + Cl_2 \longrightarrow 2\,CsCl$
   (b) $Cs \longrightarrow Cs^+ + e^-$ (oxidation) $Cl_2 + 2\,e^- \longrightarrow 2\,Cl^-$ (reduction)
   (c) Cs is the reducing agent; Cl_2 is the oxidizing agent.

3. (a) +5 (b) +3 (c) +6 (d) +6 (e) −3 (f) −1/3 (g) −3 (h) −8/3 (i) +3 (j) +5 (k) +3 (l) +7 (m) +5 (n) −2 (o) −1 (p) +3 (q) +6 (r) −4/3 (s) 0 (t) −3/2

4. a) $\underset{+4}{\mathbf{ClO_2}} + \underset{0}{\mathbf{C}} \longrightarrow \underset{+3}{\mathbf{ClO_2^-}} + \underset{+4}{\mathbf{CO_3^{2-}}}$; C is oxidized

   b) $\underset{+2}{\mathbf{Sn^{2+}}} + Cl^- + \underset{+5}{\mathbf{BrO_3^-}} \longrightarrow \underset{+4}{\mathbf{SnCl_6^{2-}}} + \underset{-1}{\mathbf{Br^-}}$; Sn is oxidized

   c) $\underset{+7}{\mathbf{MnO_4^-}} + \underset{+3}{\mathbf{C_2O_4^{2-}}} \longrightarrow \underset{+4}{\mathbf{MnO_2}} + \underset{+4}{\mathbf{CO_2}}$; C is oxidized

   d) $\underset{+5}{\mathbf{NO_3^-}} + \underset{-2}{\mathbf{H_2Te}} \longrightarrow \underset{+2}{\mathbf{NO}} + \underset{+6}{\mathbf{TeO_4^{2-}}}$; Te is oxidized

5. Equations a, d and f

6. (a) Cl_2, Cl^-, Cl_2O (b) NO_3^-, NO_2^-, N_2O_3

7. (a) Na^+ can only be reduced
   (b) I^- can only be oxidized
   (c) Cu^+ can be either reduced or oxidized
   (d) Sn^{4+} can only be reduced
   (e) NO_3^- cannot react at all (needs H^+ to react)
   (f) Hg(l) can only be oxidized
   (g) Fe^{2+} can be either reduced or oxidized
   (h) Co^{2+} can only be reduced
   (i) Se(s) cannot react at all (needs H^+ to react)
   (j) Sn^{2+} can be either reduced or oxidized
   (k) Al(s) can only be oxidized
   (l) acidic $Cr_2O_7^{2-}$ can only be reduced

8. (a) no reaction; the species to be reduced (Ni^{2+}) is below the species to be oxidized (Ag) on the Table.
   (b) spontaneous: $2\,Li + Zn^{2+} \longrightarrow 2\,Li^+ + Zn$.
   (c) no reaction; both species can only be oxidized.
   (d) no reaction; the species to be reduced (H^+) is below the species to be oxidized (Cu) on the Table.
   (e) spontaneous: $2\,H^+ + Fe \longrightarrow H_2 + Fe^{2+}$.