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Electrical Charge Method for Balancing, Quantifying, and Defining Redox Reactions

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Abstract

Defining and balancing redox reactions are core knowledge and skills in the study of chemistry. The most common method to perform these two tasks is the oxidation number method, which combines mathematical operations and application of oxidation number. However, when oxidation number is not known, it is not applicable. Algebraic methods can balance all chemical reactions mathematically, but they cannot define redox reactions chemically. This article explores the electrical charge method for balancing, quantifying, and defining redox reactions. This method only requires the balancing of atoms and electrical charges. There is no need to determine oxidation number or count the number of transferred electrons. It works effectively in complicated cases where oxidation number is uncertain and where there are more than two sets of redox couples. Furthermore, the net-charge of a redox couple can function as a counting concept to determine its number of transferred electrons and change of oxidation numbers. The electrical charge method also initiates a new charge model, which complements the conventional electron model and oxidation number model, for defining redox reactions.

Keywords: electrical charge method, ion-charge equation, net-charge, number of transferred electrons, change of oxidation numbers, redox couple, charge model

1. Introduction

Redox reactions are important in both theoretical studies and practical uses. The concept is also one of the most difficult to teach and learn (Goes, Nogueira & Fernandez, 2020). In general chemistry textbooks, the oxidation number method is a fundamental approach for counting the number of transferred electrons and understanding redox reactions (Tro, 2020; Chang & Goldsby, 2013). Without knowing oxidation number, redox reactions cannot be defined and balanced. Algebraic methods, such as linear simultaneous equations method (Porter, 1985; Olson, 1997; Kolb, 1979) and matrix method (Blakley, 1982; Risteski, 2011), can balance redox reactions, but they cannot define them chemically.

The relationships among oxidation number, transferred electrons, and electrical charge, can also be confusing for students (Garnett & Treagust, 1992; Brandriet & Bretz, 2014). In response to the limitations of the oxidation number method and the algebraic methods, the electrical charge method for balancing and defining redox reaction is developed in this article. This method does not require calculation of oxidation number nor use of electron. It only requires balancing of atoms and electrical charges by using two half reactions in a redox reaction. The key parameter is electrical charge, which acts as a concept to balance, quantify, and define redox reactions. By using simple arithmetic operations, the electrical charge method is appliable for balancing both ionic and molecular chemical equations.

2. The Electrical Charge Method

The electrical charge method is based on ion-charge equations for balancing half reactions, in which electrical charge is the key concept. There are four electrical charge parameters, which are shown as follows:

charge = individual ionic charge

Σ charge (reactant) = the sum of reactants' charge

Σ charge (product) = the sum of products' charge

net-charge = Σ charge (product) – Σ charge (reactant)

Given an example:

unbalanced ion-charge half equation: $Cr_2O_7^{2-}+H^+ \rightarrow Cr^{3+}$

balanced ion-charge half equation: $Cr_2O_7^{2-}+14H^+ \rightarrow 2Cr^{3+}+7H_2O$

(i) charge: individual ionic charge

ion	$Cr_2O_7^{2-}$	H^+	Cr^{3+}
charge	-2	+1	+3

(ii) Σ charge (reactant) and Σ charge (product)

reactants		products
$Cr_2O_7^{2-}+14H^+$	→	$2Cr^{3+}+7H_2O$
Σ charge (reactant) = 1(-2) + 14(+1) = +12		Σ charge (product) = 2(+3) = +6

(iii) net-charge = Σ charge (product) – Σ charge (reactant)

= (+6) – (+12) = -6

There are five steps in the development of the electrical charge method: (1) setting net-charge as the key parameter, (2) balancing an overall redox reaction by making its two half equations' net-charges equivalent, (3) quantifying the relationship between net-charge and number of transferred electrons, (4) establishing the charge model, and (5) defining redox reaction.

3. Procedures for Balancing Ion-Charge Equations

When using the electrical charge method, H^+, O, H_2O, and electrical charges (or charge) are employed as balancing devices. Based on the charge parameters, the balanced method is developed, and its operating procedures are shown as follows:

Step 1.	Divide the overall redox reaction into two half reactions
	(H^+, OH^-, and H_2O can be omitted in the half reactions optionally; a molecular chemical equation is converted to an ionic chemical equation when needed)
Step 2.	Balance all atoms in the two half reactions
	a) Balance all atoms except H and O
	b) Use 1 O to balance each O atom
	c) Use 1 H^+ to balance each H atom
	d) Provide 2 H^+ for each O
	e) Convert 2 H^+ and 1 O to 1 H_2O
Step 3.	Determine the net-charges of the two half reactions
	net-charge = Σ charge (product) - Σ charge (reactant)
Step 4.	Make the net-charges of the two half reactions equivalent
Step 5.	Combine the two half reactions
Step 6.	Simplify the overall chemical equation
Step 7.	Provide 1 OH^- for each H^+ and simplify the overall chemical equation
	(This is an optional step for converting an acidic solution to a basic solution.)

4. Procedures for Dividing an Overall Reaction into Two Half Reactions

The electrical charge method is a half reaction approach. The first and the most critical step is to divide an overall redox reaction into two half reactions by using the "ping-pong" strategy (Yuen & Lau, 2022a). Its working procedures are as follows: (i) choose one of the reactants and identify all its non-H and non-O elements, (ii) link the reactant's element(s) on all products' element(s), (iii) keep linking left (reactants' side)-right (products' side)-left-right…, until a half reaction is attained, (iv) choose another reactant and repeat the steps (i), (ii), and (iii).

Given an overall reaction example: $HIO_3+FeI_2+HCl \rightarrow FeCl_3+ICl+H_2O$ (Stout, 1995)

Choose the reactant HIO_3

(i) Start from $HIO_3\rightarrow$

(ii) Link from left to right: $HIO_3\rightarrow ICl$

(iii) Link from right to left: $HIO_3+HCl\rightarrow ICl$ (the first half reaction is attained)

Choose another reactant FeI_2

(i) Start from $FeI_2\rightarrow$

(ii) Link from left to right: $FeI_2\rightarrow FeCl_3+ICl$

(iii) Link from right to left: $FeI_2+HCl\rightarrow FeCl_3+ICl$ (the second half reaction is attained)

Given another overall reaction example: $CuSCN+KIO_3+HCl\rightarrow CuSO_4+KCl+HCN+ICl+H_2O$ (Stout, 1995)

Choose the reactant CuSCN

(i) Start from $CuSCN\rightarrow$

(ii) Link from left to right: $CuSCN\rightarrow CuSO_4+HCN$ (the first half reaction is attained)

Choose another reactant KIO_3

(i) Start from $KIO_3\rightarrow$

(ii) Link from left to right: $KIO_3\rightarrow KCl+ICl$

(iii) Link from right to left: $KIO_3+HCl\rightarrow KCl+ICl$ (the second half reaction is attained)

5. Examples for Balancing Redox Reactions

Example 1. In an ionic chemical equation (at acidic medium)

Given an ionic chemical equation: $CH_3CH_2OH+Cr_2O_7^{2-}+H^+\rightarrow CH_3COOH+Cr^{3+}$

Convert to $C_2H_6O+Cr_2O_7^{2-}+H^+\rightarrow C_2H_4O_2+Cr^{3+}$

Step 1.	Divide into two half reactions
	$C_2H_6O\rightarrow C_2H_4O_2$
	$Cr_2O_7^{2-}\rightarrow Cr^{3+}$
Step 2.	Balance all atoms in the two half reactions
	$C_2H_6O\rightarrow C_2H_4O_2$
	$C_2H_6O+O\rightarrow C_2H_4O_2+2H^+$
	$C_2H_6O+O+2H^+\rightarrow C_2H_4O_2+2H^++2H^+$
	$C_2H_6O+H_2O\rightarrow C_2H_4O_2+4H^+$
	$Cr_2O_7^{2-}\rightarrow Cr^{3+}$
	$Cr_2O_7^{2-}\rightarrow 2Cr^{3+}+7O$
	$Cr_2O_7^{2-}+14H^+\rightarrow 2Cr^{3+}+7O+14H^+$
	$Cr_2O_7^{2-}+14H^+\rightarrow 2Cr^{3+}+7H_2O$
Step 3.	Determine the net-charges of the two half reactions
	$C_2H_6O+H_2O\rightarrow C_2H_4O_2+4H^+$ (net-charge = +4)
	$Cr_2O_7^{2-}+14H^+\rightarrow 2Cr^{3+}+7H_2O$ (net-charge = -6)
Step 4.	Make the net-charges of the two half reactions equivalent (LCM=12)
	$(C_2H_6O+H_2O\rightarrow C_2H_4O_2+4H^+)\times 3$
	$(Cr_2O_7^{2-}+14H^+\rightarrow 2Cr^{3+}+7H_2O)\times 2$
Step 5.	Combine the two half reactions
	$3C_2H_6O+3H_2O+2Cr_2O_7^{2-}+28H^+\rightarrow 3C_2H_4O_2+12H^++4Cr^{3+}+14H_2O$
Step 6.	Simplify the overall chemical equation
	$3C_2H_6O+2Cr_2O_7^{2-}+16H^+\rightarrow 3C_2H_4O_2+4Cr^{3+}+11H_2O$

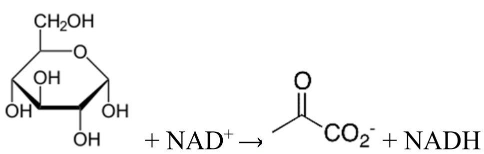
Example 2. In an ionic chemical equation (at basic medium)

Given an ionic chemical equation: $NO_2^- + MnO_4^- \rightarrow NO_3^- + MnO_2 + OH^-$

Step 1.	Divide into two half reactions	
	$NO_2^- \rightarrow NO_3^-$	
	$MnO_4^- \rightarrow MnO_2$	
Step 2.	Balance all atoms in the two half reactions	
	$NO_2^- + O \rightarrow NO_3^-$	
	$NO_2^- + O + 2H^+ \rightarrow NO_3^- + 2H^+$	
	$NO_2^- + H_2O \rightarrow NO_3^- + 2H^+$	
	$MnO_4^- \rightarrow MnO_2 + 2O$	
	$MnO_4^- + 4H^+ \rightarrow MnO_2 + 2O + 4H^+$	
	$MnO_4^- + 4H^+ \rightarrow MnO_2 + 2H_2O$	
Step 3.	Determine the net-charges of the two half reactions	
	$NO_2^- + H_2O \rightarrow NO_3^- + 2H^+$	(net-charge = +2)
	$MnO_4^- + 4H^+ \rightarrow MnO_2 + 2H_2O$	(net-charge = -3)
Step 4.	Make the net-charges of the two half reactions equivalent (LCM=6)	
	$(NO_2^- + H_2O \rightarrow NO_3^- + 2H^+) \times 3$	
	$(MnO_4^- + 4H^+ \rightarrow MnO_2 + 2H_2O) \times 2$	
Step 5.	Combine the two half reactions	
	$3NO_2^- + 3H_2O + 2MnO_4^- + 8H^+ \rightarrow 3NO_3^- + 6H^+ + 2MnO_2 + 4H_2O$	
Step 6.	Simplify the overall chemical equation	
	$3NO_2^- + 2MnO_4^- + 2H^+ \rightarrow 3NO_3^- + 2MnO_2 + H_2O$	
Step 7.	Convert the overall chemical equation from acidic solution to basic solution	
	$3NO_2^- + 2MnO_4^- + 2H^+ \rightarrow 3NO_3^- + 2MnO_2 + H_2O$	
	$3NO_2^- + 2MnO_4^- + 2H^+ + 2OH^- \rightarrow 3NO_3^- + 2MnO_2 + H_2O + 2OH^-$	
	$3NO_2^- + 2MnO_4^- + 2H_2O \rightarrow 3NO_3^- + 2MnO_2 + H_2O + 2OH^-$	
	$3NO_2^- + 2MnO_4^- + H_2O \rightarrow 3NO_3^- + 2MnO_2 + 2OH^-$	

Example 3. In an ionic chemical equation

Given a biochemical equation: glucose + $NAD^+ \rightarrow$ pyruvate + NADH



Convert to $C_6H_{12}O_6 + NAD^+ \rightarrow C_3H_3O_3^- + NADH$

Step 1.	Divide into two half reactions
	$C_6H_{12}O_6 \rightarrow C_3H_3O_3^-$
	$NAD^+ \rightarrow NADH$
Step 2.	Balance all atoms in the two half reactions
	$C_6H_{12}O_6 \rightarrow C_3H_3O_3^-$
	$C_6H_{12}O_6 \rightarrow 2C_3H_3O_3^-$
	$C_6H_{12}O_6 \rightarrow 2C_3H_3O_3^- + 6H^+$
	$NAD^+ \rightarrow NADH$
	$NAD^+ + H^+ \rightarrow NADH$
Step 3.	Determine the net-charges of the two half reactions
	$C_6H_{12}O_6 \rightarrow 2C_3H_3O_3^- + 6H^+$ (net-charge = +4)
	$NAD^+ + H^+ \rightarrow NADH$ (net-charge = -2)
Step 4.	Make the net-charges of the two half reactions equivalent (LCM=4)
	$(C_6H_{12}O_6 \rightarrow 2C_3H_3O_3^- + 6H^+) \times 1$
	$(NAD^+ + H^+ \rightarrow NADH) \times 2$
Step 5.	Combine the two half reactions
	$C_6H_{12}O_6 + 2NAD^+ + 2H^+ \rightarrow 2C_3H_3O_3^- + 6H^+ + 2NADH$
Step 6.	Simplify the overall chemical equation
	$C_6H_{12}O_6 + 2NAD^+ \rightarrow 2C_3H_3O_3^- + 2NADH + 4H^+$

Example 4. In a molecular chemical equation containing multiple sets of redox couples

Given an inorganic equation: $Pb(N_3)_2 + Cr(MnO_4)_2 \rightarrow Pb_3O_4 + NO + Cr_2O_3 + MnO_2$

Step 1.	Divide into two half reactions
	$Pb(N_3)_2 \rightarrow Pb_3O_4 + NO$
	$Cr(MnO_4)_2 \rightarrow Cr_2O_3 + MnO_2$
Step 2.	Balance all atoms in the two half reactions
	$Pb(N_3)_2 \rightarrow Pb_3O_4 + NO$
	$3Pb(N_3)_2 \rightarrow Pb_3O_4 + 18NO$
	$3Pb(N_3)_2 + 22O \rightarrow Pb_3O_4 + 18NO$
	$3Pb(N_3)_2 + 22O + 44H^+ \rightarrow Pb_3O_4 + 18NO + 44H^+$
	$3Pb(N_3)_2 + 22H_2O \rightarrow Pb_3O_4 + 18NO + 44H^+$
	$Cr(MnO_4)_2 \rightarrow Cr_2O_3 + MnO_2$
	$2Cr(MnO_4)_2 \rightarrow Cr_2O_3 + 4MnO_2$
	$2Cr(MnO_4)_2 \rightarrow Cr_2O_3 + 4MnO_2 + 5O$
	$2Cr(MnO_4)_2 + 10H^+ \rightarrow Cr_2O_3 + 4MnO_2 + 5O + 10H^+$
	$2Cr(MnO_4)_2 + 10H^+ \rightarrow Cr_2O_3 + 4MnO_2 + 5H_2O$
Step 3.	Determine the net-charges of the two half reactions
	$3Pb(N_3)_2 + 22H_2O \rightarrow Pb_3O_4 + 18NO + 44H^+$ (net-charge = +44)
	$2Cr(MnO_4)_2 + 10H^+ \rightarrow Cr_2O_3 + 4MnO_2 + 5H_2O$ (net-charge = -10)
Step 4.	Make the net-charges of the two half reactions equivalent (LCM=220)
	$(3Pb(N_3)_2 + 22H_2O \rightarrow Pb_3O_4 + 18NO + 44H^+) \times 5$
	$(2Cr(MnO_4)_2 + 10H^+ \rightarrow Cr_2O_3 + 4MnO_2 + 5H_2O) \times 22$
Step 5.	Combine the two half reactions

$15Pb(N_3)_2+110H_2O+44Cr(MnO_4)_2+220H^+ \rightarrow 5Pb_3O_4+90NO+220H^++$
$22Cr_2O_3+88MnO_2+110H_2O$

Step 6. Simplify the overall chemical equation

$15Pb(N_3)_2+44Cr(MnO_4)_2 \rightarrow 5Pb_3O_4+90NO+22Cr_2O_3+88MnO_2$

Example 5. In a molecular chemical equation containing uncertain oxidation numbers

Given a molecular chemical equation: $Fe_7C_3+HNO_3 \rightarrow Fe(NO_3)_3+CO_2+NO_2+H_2O$

Step 1. Divide into two half reactions

$Fe_7C_3+HNO_3 \rightarrow Fe(NO_3)_3+CO_2$

$HNO_3 \rightarrow NO_2$

Step 2. Balance all atoms in the two half reactions

$Fe_7C_3+HNO_3 \rightarrow Fe(NO_3)_3+CO_2$

$Fe_7C_3+HNO_3 \rightarrow 7Fe(NO_3)_3+3CO_2$

$Fe_7C_3+21HNO_3 \rightarrow 7Fe(NO_3)_3+3CO_2$

$Fe_7C_3+21HNO_3+6O \rightarrow 7Fe(NO_3)_3+3CO_2+21H^+$

$Fe_7C_3+21HNO_3+6O+12H^+ \rightarrow 7Fe(NO_3)_3+3CO_2+21H^++12H^+$

$Fe_7C_3+21HNO_3+6H_2O \rightarrow 7Fe(NO_3)_3+3CO_2+33H^+$

$HNO_3 \rightarrow NO_2$

$HNO_3 \rightarrow NO_2+O+H^+$

$HNO_3+H^+ \rightarrow NO_2+O+H^++H^+$

$HNO_3+H^+ \rightarrow NO_2+H_2O$

Step 3. Determine the net-charges of the two half reactions

$Fe_7C_3+21HNO_3+6H_2O \rightarrow 7Fe(NO_3)_3+3CO_2+33H^+$ (net-charge = +33)

$HNO_3+H^+ \rightarrow NO_2+H_2O$ (net-charge = -1)

Step 4. Make the net-charges of the two half reactions equivalent (LCM=33)

$(Fe_7C_3+21HNO_3+6H_2O \rightarrow 7Fe(NO_3)_3+3CO_2+33H^+) \times 1$

$(HNO_3+H^+ \rightarrow NO_2+H_2O) \times 33$

Step 5. Combine the two half reactions

$Fe_7C_3+21HNO_3+6H_2O+33HNO_3+33H^+ \rightarrow 7Fe(NO_3)_3+3CO_2+33H^++33NO_2+33H_2O$

Step 6. Simplify the overall chemical equation

$Fe_7C_3+54HNO_3 \rightarrow 7Fe(NO_3)_3+3CO_2+33NO_2+27H_2O$

Example 6. In a complicated chemical equation

Given a molecular chemical equation (Stout, 1995; Ludwig, 1996; Hart, 1996; Nelson, 1997; Herndon, 1997)

$[Cr(N_2H_4CO)_6]_4[Cr(CN)_6]_3+KMnO_4+H_2SO_4 \rightarrow K_2Cr_2O_7+MnSO_4+CO_2+KNO_3+K_2SO_4+H_2O$

Convert to $Cr_7N_{66}H_{96}C_{42}O_{24}+KMnO_4+H_2SO_4 \rightarrow K_2Cr_2O_7+MnSO_4+CO_2+KNO_3+K_2SO_4+H_2O$

Convert to an ionic chemical equation (to eliminate spectator ions K^+; SO_4^{2-})

$Cr_7N_{66}H_{96}C_{42}O_{24}+MnO_4^-+H^+ \rightarrow Cr_2O_7^{2-}+Mn^{2+}+CO_2+NO_3^-+H_2O$

Step 1. Divide into two half reactions

$Cr_7N_{66}H_{96}C_{42}O_{24} \rightarrow Cr_2O_7^{2-}+CO_2+NO_3^-$

$MnO_4^- \rightarrow Mn^{2+}$

Step 2. Balance all atoms in the two half reactions

$Cr_7N_{66}H_{96}C_{42}O_{24} \rightarrow Cr_2O_7^{2-}+CO_2+NO_3^-$

$2Cr_7N_{66}H_{96}C_{42}O_{24} \rightarrow Cr_2O_7^{2-}+CO_2+NO_3^-$

$2Cr_7N_{66}H_{96}C_{42}O_{24}+565O \rightarrow 7Cr_2O_7^{2-}+84CO_2+132NO_3^-+192H^+$

$2Cr_7N_{66}H_{96}C_{42}O_{24}+565O+1130H^+ \rightarrow 7Cr_2O_7^{2-}+84CO_2+132NO_3^-+192H^++1130H^+$

$2Cr_7N_{66}H_{96}C_{42}O_{24}+565H_2O \rightarrow 7Cr_2O_7^{2-}+84CO_2+132NO_3^-+1322H^+$

$MnO_4^- \rightarrow Mn^{2+}+4O$

$MnO_4^-+8H^+ \rightarrow Mn^{2+}+4O+8H^+$

$MnO_4^-+8H^+ \rightarrow Mn^{2+}+4H_2O$

Step 3. Determine the net-charges of the two half reactions

$2Cr_7N_{66}H_{96}C_{42}O_{24}+565H_2O \rightarrow 7Cr_2O_7^{2-}+84CO_2+132NO_3^-+1322H^+$

(Net charge = +1176)

$MnO_4^-+8H^+ \rightarrow Mn^{2+}+4H_2O$ (Net charge = -5)

Step 4. Make the net-charges of the two half reactions equivalent (LCM=5880)

$(2Cr_7N_{66}H_{96}C_{42}O_{24}+565H_2O \rightarrow 7Cr_2O_7^{2-}+84CO_2+132NO_3^-+1322H^+) \times 5$

$(MnO_4^-+8H^+ \rightarrow Mn^{2+}+4H_2O) \times 1176$

Step 5. Combine the two half reactions

$10Cr_7N_{66}H_{96}C_{42}O_{24}+2825H_2O+1176MnO_4^-+9408H^+$

$\rightarrow 35Cr_2O_7^{2-}+420CO_2+660NO_3^-+6610H^++1176Mn^{2+}+4704H_2O$

Step 6. Simplify the overall ionic equation and convert to the overall molecular equation

$10Cr_7N_{66}H_{96}C_{42}O_{24}+1176MnO_4^-+2798H^+ \rightarrow 35Cr_2O_7^{2-}+1176Mn^{2+}+420CO_2+660NO_3^-+1879H_2O$

$10Cr_7N_{66}H_{96}C_{42}O_{24}+1176MnO_4^-+1176K^++1399SO_4^{-2}+2798H^+$
$\rightarrow 35Cr_2O_7^{2-}+70K^++1176Mn^{2+}+1176SO_4^{-2}+420CO_2+660NO_3^-+660K^++1879H_2O+223SO_4^{-2}+446K^+$

$10[Cr(N_2H_4CO)_6]_4[Cr(CN)_6]_3+1176KMnO_4+1399H_2SO_4$

$\rightarrow 35K_2Cr_2O_7+1176MnSO_4+420CO_2+660KNO_3+223K_2SO_4+1879H_2O$

6. Relationships among Net-Charge, Number of Transferred Electrons, and Change of Oxidation Numbers

6.1 Net-charge and Number of Transferred Electrons

The nature of redox reaction is an electron-transfer reaction. An ion-charge equation can be converted to an ion-electron equation by adding electrons to make the number of electrical charges even on both reactants' side and products' side. The quantitative relationship between net-charge and number of transferred electrons (Te^-) in a half reaction is demonstrated in the following examples.

In Example 1, the conversion is shown below:

ion-charge equation: $Cr_2O_7^{2-}+14H^+ \rightarrow 2Cr^{3+}+7H_2O$

net-charge = Σ charge (product) - Σ charge (reactant)

net-charge = (+6) – (+12) = -6

ion-electron equation: $Cr_2O_7^{2-}+14H^++6e^-\rightarrow 2Cr^{3+}+7H_2O$

Te^- (6 electrons on the reactants' side) = -6

Te^- (gain of 6 electrons) = -6

net-charge = Te^-

In Example 2 of "$NO_2^-+MnO_4^-\rightarrow NO_3^-+MnO_2+OH^-$", the half reaction of "$NO_2^-+2OH^-\rightarrow NO_3^-+2H_2O$" has a net-charge of +2, which represents a loss of 2 electrons ("$NO_2^-+2OH^-\rightarrow NO_3^-+2H_2O+2e^-$"). The other half reaction of "$MnO_4^-+2H_2O\rightarrow MnO_2+4OH^-$" has a net-charge of -3, which represents a gain of 3 electrons ("$MnO_4^-+2H_2O+3e^-\rightarrow MnO_2+4OH^-$").

6.2 Net-charge and Change of Oxidation Numbers

The mathematical equation of Te^- = n ΔON among Te^-, number of atoms with oxidation numbers change (n), and change of oxidation numbers (ΔON) for one set of redox couple has been established (Yuen & Lau, 2022b). In this article, the quantitative relationship between net-charge, n, and ΔON in a half reaction is shown as:

net-charge = Σ charge (product) - Σ charge (reactant)

net charge = Te^-

Te^- = n ΔON

net charge = n ΔON

By using Example 1 as a demonstration, CH_3CH_2OH/CH_3COOH, $Cr_2O_7^{2-}/Cr^{3+}$ are two individual sets of redox couples existing in two half reactions respectively.

net-charge = Σ charge (product) - Σ charge (reactant)

$CH_3CH_2OH+H_2O\rightarrow CH_3COOH+4H^+$

net-charge = {(+4) – (0)} = +4

$Cr_2O_7^{2-}+14H^+\rightarrow 2Cr^{3+}+7H_2O$

net-charge = {(+6) – (+12)} = -6

net charge = n ΔON

ΔON = ON (product) - ON (reactant)

net charge = n {ON (product) - ON (reactant)}

net charge = n_C ΔON_C

net charge = n_C {ON_C (CH_3COOH) – ON_C (CH_3CH_2OH)}

net charge = (2) {(0) – (–2)} = (2) (+2) = +4

net charge = n_{Cr} ΔON_{Cr}

net charge = n_{Cr} {ON_{Cr} (Cr^{3+}) – ON_{Cr} ($Cr_2O_7^{2-}$)}

net charge = (2) {(+3) – (+6)} = (2) (–3) = –6

When the net charges = +4, it represents an increase of oxidation number of two carbon atoms (ΔON_C = +2; n_C = 2) from CH_3CH_2OH (ON_C = -2) to CH_3COOH (ON_C = 0) whereas when net-charge = -6, it represents a decrease of oxidation number of two chromium atoms (ΔON_{Cr} = -3; n_{Cr} = 2) from $Cr_2O_7^{2-}$ (ON_{Cr} = +6) to $2Cr^{3+}$ (ON_{Cr} = +3).

6.3 Triangular Relationships among Net-charge, Te^-, and ΔON

All half redox reactions in Examples 1 to 3 contain one set of redox couple. The net-charge or Te^- of a redox couple can be counted by their ON (reactant), ON (product), and n. Their redox natures are shown in Table 1. The triangular relationships among net-charge, Te^-, and ΔON are demonstrated in Figure 1.

net charge = n ΔON = n {ON (product) - ON (reactant)}

Te^- = n ΔON = n {ON (product) - ON (reactant)}

Table 1. Net-charge, Te^-, and ΔON for defining half redox reactions

Redox terms	Net-charge	Te^-	ΔON
oxidation	net-charge increase net-charge > 0 (+) value	loss of electrons $Te^- > 0$ (+) value	oxidation number increase $\Delta ON > 0$ (+) value
reduction	net-charge decrease net-charge < 0 (-) value	gain of electrons $Te^- < 0$ (-) value	oxidation number decrease $\Delta ON < 0$ (-) value

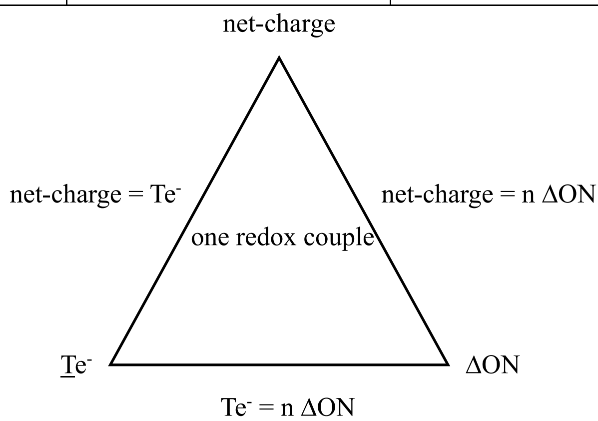


Figure 1. Triangular relationships among net-charge, Te^-, and ΔON

With reference to Table 1 and Figure 1, the selected half reactions in Table 2 can be quantified, classified, and defined.

Table 2. Net-charge, Te^-, n, and ΔON for quantifying and defining half redox reactions containing one set of redox couple

Balanced half ion-charge equation	Balanced half ion-electron equation	Net-charge	Te^-	n	ΔON	Type of reaction
$NO_3^-+4H^+ \rightarrow NO+2H_2O$	$NO_3^-+4H^++3e^- \rightarrow NO+2H_2O$	-3	-3	1	-3	reduction
$CO_3^{2-}+2H^+ \rightarrow CO_2+2H_2O$	$CO_3^{2-}+2H^+ \rightarrow CO_2+2H_2O$	0	0	1	0	non-redox
$C_2H_6O+H_2O \rightarrow C_2H_4O_2+4H^+$	$C_2H_6O+H_2O \rightarrow C_2H_4O_2+4H^++4e^-$	+4	+4	2	+2	oxidation
$C_6H_{12}O_6 \rightarrow 2C_3H_3O_3^-+6H^+$	$C_6H_{12}O_6 \rightarrow 2C_3H_3O_3^-+6H^++4e^-$	+4	+4	6	$+\frac{2}{3}$	oxidation
$Cr_2O_7^{2-}+14H^+ \rightarrow 2Cr^{3+}+7H_2O$	$Cr_2O_7^{2-}+14H^++6e^- \rightarrow 2Cr^{3+}+7H_2O$	-6	-6	2	-3	reduction
$MnO_4^-+4H^+ \rightarrow MnO_2+2H_2O$	$MnO_4^-+4H^++3e^- \rightarrow MnO_2+2H_2O$	-3	-3	2	-3	reduction

The relationships among the loss/gain of electrons, the increase/decrease of oxidation number, and the increase/decrease of charge are established in a half redox reaction. Regarding their triangular relationships, an example of oxidation is shown in Figure 2.

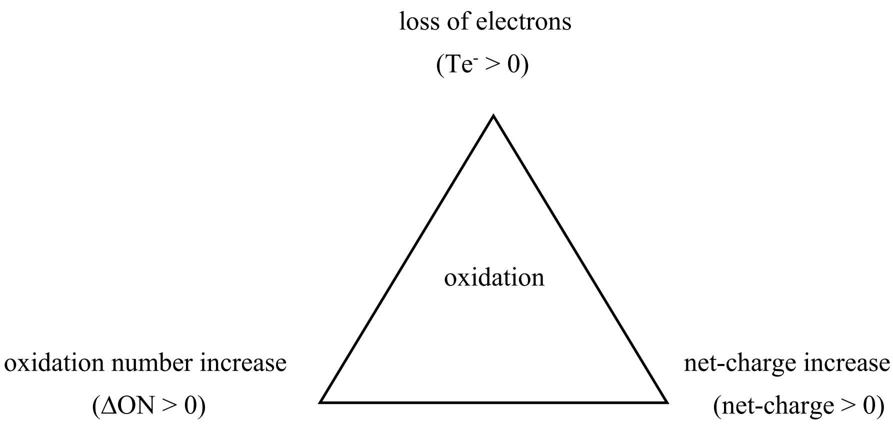


Figure 2. Triangular relationships among number of transferred electrons, change of oxidation numbers, and net-charge in a half oxidation reaction

6.4 Multiple Redox Couples in a Half Redox Reaction

By using Example 4 as a demonstration, the overall reaction of "$Pb(N_3)_2+Cr(MnO_4)_2 \rightarrow Pb_3O_4+NO+Cr_2O_3+MnO_2$" contains four sets of redox couples. In the first half oxidation reaction of "$3Pb(N_3)_2+22H_2O \rightarrow Pb_3O_4+18NO+44H^+$", the resulting net-charge is +44 (Te^- = +44; a loss of 44 electrons). In the second half reduction reaction of "$2Cr(MnO_4)_2+10H^+ \rightarrow Cr_2O_3+4MnO_2+5H_2O$", the resulting net charge equals -10 (Te^- = -10; a gain of 10 electrons).

In the half oxidation reaction, there are two sets of redox couples shown as Pb^{2+}/Pb_3O_4 and N_3^-/NO. The calculations of net-charges for them are shown as follows:

For $3Pb^{2+} \rightarrow Pb_3O_4$

$$\text{net charge} = n_{Pb}\ \Delta ON_{Pb}$$

$$\text{net charge} = n_{Pb}\ \{ON_{Pb}(Pb_3O_4) - ON_{Pb}(Pb^{2+})\}$$

$$\text{net charge} = (3)\ \{(+\tfrac{8}{3}) - (+2)\} = (3)\ (+\tfrac{2}{3})) = +2$$

For $6N_3^- \rightarrow 18NO$

$$\text{net charge} = n_N\ \Delta ON_N$$

$$\text{net charge} = n_N\ \{ON_N(NO) - ON_N(N_3^-)\}$$

$$\text{net charge} = (18)\ \{(+2) - (-\tfrac{1}{3})\} = (18)\ (+\tfrac{7}{3}) = +42$$

$$\Sigma \text{ net-charge (redox couple)} = \text{net-charge } (3Pb^{2+} \rightarrow Pb_3O_4) + \text{net-charge } (6N_3^- \rightarrow 18NO)$$

$$= (+2) + (+42)$$

$$= +44$$

In another half reduction reaction, there are two sets of redox couples shown as Cr^{2+}/Cr_2O_3 and MnO_4^-/MnO_2. The calculations of net-charges for them are shown as follows:

For $2Cr^{2+} \rightarrow Cr_2O_3$

$$\text{net charge} = n_{Cr}\ \Delta ON_{Cr}$$

$$\text{net charge} = n_{Cr}\ \{ON_{Cr}(Cr_2O_3) - ON_{Cr}(Cr^{2+})\}$$

$$\text{net charge} = (2)\ \{(+3) - (+2)\} = (2)\ (+1) = +2$$

For $4MnO_4^- \rightarrow 4MnO_2$

$$\text{net charge} = n_{Mn}\ \Delta ON_{Mn}$$

$$\text{net charge} = n_{Mn}\{ON_{Mn}(MnO_2) - ON_{Mn}(MnO_4^-)\}$$

$$\text{net charge} = (4)\ \{(+4) - (+7)\} = (4)\ (-3) = -12$$

$$\Sigma \text{ net-charge (redox couple)} = \text{net-charge } (2Cr^{2+} \rightarrow Cr_2O_3) + \text{net-charge } (4MnO_4^- \rightarrow 4MnO_2)$$

$$= (+2) + (-12)$$

$$= -10$$

The net-charge, n, and ΔON for multiple redox couples in Example 4 are summarized in Table 3. The mathematical relationships between a half reaction and multiple redox couples are demonstrated as follows:

$$\text{net-charge (half reaction)} = \Sigma \text{ net-charge (redox couple)}$$

$$\text{net-charge (half reaction)} = \Sigma\ n\ \Delta ON \text{ (redox couple)}$$

$$Te^- \text{ (half reaction)} = \Sigma\ Te^- \text{ (redox couple)}$$

$$Te^- \text{ (half reaction)} = \Sigma\ n\ \Delta ON \text{ (redox couple)}$$

Table 3. Relationships among net-charge, n, and ΔON for multiple redox couples in a balanced redox reaction

Half reaction	Net-charge	Redox type	Redox couple	Net-charge	n	ΔON	Redox type
$3Pb(N_3)_2+22H_2O \rightarrow Pb_3O_4+18NO+44H^+$	+44	oxidation	$3Pb^{2+} \rightarrow Pb_3O_4$	+2	3	$+\frac{2}{3}$	oxidation
			$6N_3^- \rightarrow 18NO$	+42	18	$+\frac{7}{3}$	oxidation
$2Cr(MnO_4)_2+10H^+ \rightarrow Cr_2O_3+4MnO_2+5H_2O$	–10	reduction	$2Cr^{2+} \rightarrow Cr_2O_3$	+2	2	+1	oxidation
			$4MnO_4^- \rightarrow 4MnO_2$	–12	4	–3	reduction

With reference to Table 3, the half reduction reaction of “$2Cr(MnO_4)_2+10H^+ \rightarrow Cr_2O_3+4MnO_2+5H_2O$” is quantified and defined by its net-charge of –10 (reduction), or summation of net-charges of Cr^2/Cr_2O_3 (+2; oxidation) and MnO_4/MnO_2 (–12; reduction).

7. The Charge Model: A New Redox Model

The establishment of the electrical charge method for balancing redox reactions initiates and generalizes a new charge model. A comparison of the electron (e^-) model, the oxidation number (ON) model (IUPAC, 2019), and the charge model is shown in Table 4.

Table 4. Comparison of the electron model, the oxidation number model, and the charge model for redox reactions

Redox terms	Electron model	Oxidation number model	Charge model
oxidation	loss of e^-	ON increase	charge increase
reduction	gain of e^-	ON decrease	charge decrease
oxidizing agent	gain of e^-	ON decrease	charge decrease
reducing agent	loss of e^-	ON increase	charge increase

Net-charge can be a redox concept, which indicates charge increase or decrease in any balanced half ion-charge equation. With reference to Table 4, all ion-charge equations in Examples 1 – 6 can be defined by the charge model and summarized in Table 5.

Table 5. Charge model for defining half redox reactions

Balanced ion-charge half equation	Charge	Type of redox reaction
$C_2H_6O+H_2O \rightarrow C_2H_4O_2+4H^+$	+4; increase	oxidation
$Cr_2O_7^{2-}+14H^+ \rightarrow 2Cr^{3+}+7H_2O$	-6; decrease	reduction
$NO_2^-+H_2O \rightarrow NO_3^-+2H^+$	+2; increase	oxidation
$MnO_4^-+4H^+ \rightarrow MnO_2+2H_2O$	-3; decrease	reduction
$C_6H_{12}O_6 \rightarrow 2C_3H_3O_3^-+6H^+$	+4; increase	oxidation
$NAD^++H^+ \rightarrow NADH$	-2; decrease	reduction
$3Pb(N_3)_2+22H_2O \rightarrow Pb_3O_4+18NO+44H^+$	+44; increase	oxidation
$2Cr(MnO_4)_2+10H^+ \rightarrow Cr_2O_3+4MnO_2+5H_2O$	–10; decrease	reduction
$Fe_7C_3+21HNO_3+6H_2O \rightarrow 7Fe(NO_3)_3+3CO_2+33H^+$	+33; increase	oxidation
$HNO_3+H^+ \rightarrow NO_2+H_2O$	-1; decrease	reduction
$2Cr_7N_{66}H_{96}C_{42}O_{24}+565H_2O \rightarrow 7Cr_2O_7^{2-}+84CO_2+132NO_3^-+1322H^+$	+1176; increase	oxidation
$MnO_4^-+8H^+ \rightarrow Mn^{2+}+4H_2O$	-5; decrease	reduction

8. Conclusion

The misconceptions among oxidation number, transferred electron, and electrical charge can cause difficulties for understanding redox reactions. This article studies the electrical charge method. The significance of this method is that it manifests both chemical and mathematical accuracy. It only requires simple mathematical manipulations for balancing, and then defining redox reactions. Instead of using the oxidation number and electron, it balances ion-charge equations and counts electrical charges. Due to this functionality, it works well for half redox reactions, including complicated cases where oxidation numbers are uncertain and where there are more than two sets of redox couples. Furthermore, the resulting net-charge can be applied to determine the number of transferred electrons in any balanced half redox equation. The electrical charge method also initiates and establishes a new charge model, which complements the conventional electron model and the oxidation number model for defining redox reactions.

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