

# Balancing Redox Chemical Equations: A Discovery Procedure Employing Oxidation Reduction Titration

S. Ghaffari\*, P. K. Thamburaj, S. Abu-Baker, Annette Holstein

Chemistry Department, Ohio University Zanesville, Zanesville Ohio, United States

**Abstract** The relationship between an oxidizing agent and reducing agent may be established by a volumetric procedure known as redox titration. Potassium permanganate,  $KMnO_4$ , is a favorite oxidant partly because of its color which serves as the indicator. Iron (II) ion,  $Fe^{2+}$ , as reducing agent is titrated with  $KMnO_4$  to determine the oxidation state of  $Mn$  in  $KMnO_4$ . In the second titration Oxalate ion,  $C_2O_4^{2-}$ , is used as a reducing agent against the  $KMnO_4$ . Data obtained from titrations will lead to balancing each redox chemical reaction in an acidic medium.

**Keywords** Balancing Chemical Equation, Oxidation/Reduction, Titration/Volumetric Analysis, First-Year Undergraduate/General Chemistry

## 1. Introduction

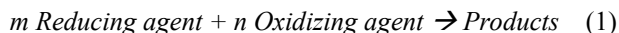
A chemical equation is a symbolic representation of a chemical reaction. To satisfy the law of conservation of mass (matter) these equations must be balanced. Balanced chemical equations are essential to solve problems in stoichiometry.

In essence, balancing chemical equations is a mathematical procedure. Most chemical equations may be balanced by a simple trial and error or inspection technique. However, for chemical reactions labeled as redox reactions, it may seem that there is no simple method for balancing.

There is a large number of articles published dealing with a variety of redox reactions [1-13]. These range from inspection to algebraic method.

In redox equations, the number of electrons transferred from a reducing agent (oxidized substance) to an oxidizing agent (reduced substance) has to be balanced as well.

In general, a redox equation can be expressed as



In this equation “ $m$ ” is the number of electrons lost by a single unit of a reducing agent and “ $n$ ” is the number of electrons gained by a single unit of an oxidizing agent. In redox reactions, the amount associated with one mole of electron change is termed as one gram-equivalent and so,

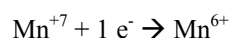
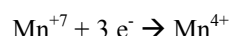
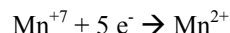
*No. of gram-equivalents of reducing agent = No. of gram-equivalents of oxidizing agent*

*$m \times \text{moles of reducing agent} = n \times \text{moles of oxidizing agent}$*

$$\frac{m}{n} = \frac{\text{moles of reducing agent}}{\text{moles of oxidizing agent}} \quad (2)$$

Equation 2, clearly implies the inverse relationship between electrons transferred and moles of the substance used in a process.

The manganese in  $KMnO_4$  has an oxidation state of “+7” and there are three possible reductions for  $Mn^{+7}$  to lower oxidation states. The half-reactions are presented below,

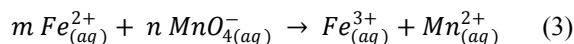


To find the ratio of “ $m/n$ ”, molarity and the volume of reducing agent solution used in titration and molarity of  $MnO_4^-$  solution are provided. The volume of  $MnO_4^-$  required to reach the equivalence point is obtained by a titration experiment.

Using a redox titration approach to find electron change of the oxidizing agent,  $KMnO_4$ , and balancing redox equations involving  $KMnO_4$  at an introductory chemistry level is described here.

## 2. Methodology

**Part I.  $Fe^{2+}/MnO_4^-$ :** The oxidation reduction reaction of  $Fe^{2+}$  with  $KMnO_4$  is presented by an unbalanced equation with “ $m$ ” and “ $n$ ” as coefficients of  $Fe^{2+}$  ion and  $MnO_4^-$  ion. These coefficients balance the number of electrons transferred from a reducing agent to an oxidizing agent.



Number of moles is given by “*molarity (M) x volume (V).*”

Substituting “ $MV$ ” for the number of moles in Eqn. 2 produces Eqn. 4. This can be used for all redox reactions.

\* Corresponding author:

ghaffari@ohio.edu (S. Ghaffari)

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$$\frac{m}{n} = \frac{M_R V_R}{M_O V_O} \quad (4)$$

Subscripts “R” and “O” represent reducing and oxidizing agents respectively.

In Eqn. 4,  $Fe^{2+}$  is the substitute for a reducing agent and  $MnO_4^-$  for the oxidizing agent.

$$\frac{m}{n} = \frac{M_{Fe^{2+}} V_{Fe^{2+}}}{M_{MnO_4^-} V_{MnO_4^-}}$$

The method of finding the ratio of “m/n” molarity and the volume of  $Fe^{2+}$  solution and molarity of  $MnO_4^-$  solution are provided. The volume of  $MnO_4^-$  required to reach the equivalence point is obtained by titration.

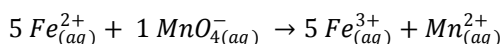
Student-generated experimental results give

$$\frac{m}{n} = 4.90 \pm 0.46$$

This will be presented as the ratio of two whole numbers

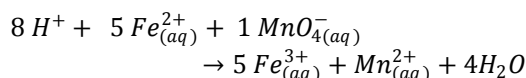
$$\frac{m}{n} = \frac{5}{1}$$

“m” and “n” in Eqn 3 are replaced by their values “5” and “1” respectively and coefficients of the product side are added accordingly to balance elements that are oxidized “Fe” and reduced “Mn”.

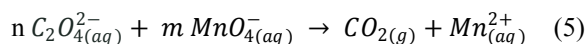


Iron has only two oxidation states “+2” and “+3” and there are five electrons exchanged between the reducing agent and the oxidizing agent. Therefore, Mn in  $KMnO_4$  must have an oxidation state of “+7”.

To complete the balancing of the reaction, oxygen and hydrogen must be balanced, too. To balance the number of oxygens for each oxygen needed, one molecule of  $H_2O$  is added. Since this reaction is in acidic medium, to balance hydrogen  $H^+$ 's are added to complete the balancing of the redox reaction.



**Part II.  $C_2O_4^{2-}/MnO_4^-$ :** In the second titration oxalate ion,  $C_2O_4^{2-}$ , is used as the reducing agent in titration with  $KMnO_4$ .



In Eqn. 4,  $C_2O_4^{2-}$  is a substitute for a reducing agent and  $MnO_4^-$  for the oxidizing agent.

$$\frac{m}{n} = \frac{M_{C_2O_4^{2-}} V_{C_2O_4^{2-}}}{M_{MnO_4^-} V_{MnO_4^-}}$$

Student-generated experimental results give

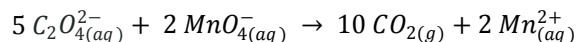
$$\frac{m}{n} = 2.51 \pm 0.06$$

$$\frac{m}{n} = \frac{2.51}{1}$$

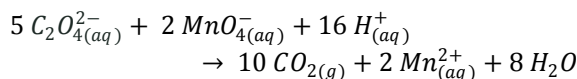
To round this ratio to a ratio of two whole numbers gives

$$\frac{m}{n} = \frac{5}{2}$$

“m” and “n” in Eqn 5 are replaced by their values “5” and “2” respectively and coefficients of the product side are added accordingly.



To complete balancing the reaction in an acidic medium, oxygen and hydrogen must be balanced following the procedure described previously. The final balanced equation is,



### 3. Conclusions

This method is student friendly and provides hands-on experience and confirmation of the algebraic method of balancing redox reactions.

Further study is needed regarding the application of this method in organic redox reactions and the use of other oxidizing agents such as dichromate.

### REFERENCES

- [1] Swinehart, D. F., More on chemical reaction balancing, *J. Chem. Educ.* 1985, 62(1), 55.
- [2] Blakley, G. R., Chemical equation balancing: A general method which is quick, simple, and has unexpected applications, *J. Chem. Educ.* 1982, 59(9), 728.
- [3] Olson, J. A., An Analysis of the Algebraic Method for Balancing Chemical Reactions, *J. Chem. Educ.* 1997, 74(5), 538-42.
- [4] Harjadi, W. J., A simpler method of chemical reaction balancing, *J. Chem. Educ.* 1986, 63(11), 978-9.
- [5] Kolb, D., More On Balancing Redox Equations, *J. Chem. Educ.* 1979, 56(3), 181-4.
- [6] Hoor, M. J., Redox Balancing without Puzzling, *J. Chem. Educ.* 1997, 74(11), 1367-8.
- [7] Burrell, H. P. C., Balancing organic redox equations, *J. Chem. Educ.* 1959, 36(2), 77-9.
- [8] Jensen, W. B., Balancing Redox Reactions, *J. Chem. Educ.*, 2009, 86(6), 681-2
- [9] Pernicone, E., Method for balancing redox reactions containing hydroxyl ions, *J. Chem. Educ.*, 1981, 58, 966.
- [10] Ludwing, O. G., On Balancing “Redox Challenges”, *J. Chem. Educ.*, 1996, 73(6), 507.
- [11] Carrano, S. A., Balancing an Atypical Redox Equation, *J. Chem. Educ.*, 1978, 55(6), 382.
- [12] Kolb, D., Balancing Complex Redox Equation by Inspection, *J. Chem. Educ.*, 1981, 58(8), 642-5.

- [13] Ludwing, O. G., On Balancing "Redox Challenges" by Unconventional Oxidation Numbers, *J. Chem. Educ.*, 1997, 73(11), 1256-7.