

Nonequilibrium thermodynamics of energy transformations

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Abstract

Important aspects of classical and nonequilibrium thermodynamics developed separately and lack alignment. Confusing definitions of heat and work exist for open systems in association with the exchange of matter. The concept of *heat* exchange in open systems requires clarification on the basis of calorimetric principles, whereas *power* is rigorously defined in terms of external work per time and the product of internal flows and forces. For illustration, analogous electric, thermal and chemical flows and forces are represented. An internal *flow* is the *advancement* of a transformation per time. A *force* is the partial Gibbs (Helmholtz) energy change per advancement. These relations are developed on the basis of the second law of thermodynamics with reference to entropy production, *efficiency and energy dissipation*. The symbols of nonequilibrium thermodynamics are not generally in line with IUPAC conventions. Any attempt towards a reconciliation necessarily leads to symbols which are unconventional in either tradition. Importantly, improvement of terminological consistency is a basis for conceptual clarification.

INTRODUCTION

Advancement in the understanding of biological energy transformations relies upon the integration of a wide spectrum of concepts in physical chemistry. Interdisciplinary communication requires a standard notation in this expanding field of biophysical chemistry. Attempts towards a concise terminology are aided by IUPAC recommendations (ref. 1). However, alignment of nomenclature in nonequilibrium thermodynamics, kinetics, physiological energetics, bioenergetics, clinical bioenergetics and biotechnology is complicated by the fact that several areas developed their special terminology with limited interconnection. When crossing the arbitrary borders between these disciplines, one is faced with contrasting conventions regarding nomenclature, units and symbols. Unification of terminology and strict application of *SI* units, therefore, would facilitate communication and collaboration. Moreover, it would lead to a clarification of concepts on irreversible processes in complex open systems (ref. 2). This paper is an attempt to formulate a coherent nomenclature incorporating IUPAC guidelines into nonequilibrium thermodynamics. In so doing it became obvious that aspects of traditional terminology require revision. The following considerations are not a recommendation but are conceived as the author's contribution to the substantial effort required for developing a uniform nomenclature for classical and nonequilibrium thermodynamics.

The list of symbols in the Appendix summarizes some fundamental physico-chemical quantities for describing energy transformations. Notably, the term energy transformation is general and vague, since energy may be associated with either the first or second law of thermodynamics. Generalization of flows and forces extends the unifying concepts of energy and power related to entropy production. *Extensive* quantities (Table A1) pertain to a total system, whereas *specific* quantities relate to the unit size of the system (Table A2). It is suggested to separate the specific from the *intensive* quantities, the latter including the forces (Tables A1 and A2). Tables A3 to A5 provide examples for analogous descriptions of electric, thermal and chemical processes. Table A6 summarizes flux and force ratios, related to stoichiometry and efficiency as relevant for biological calorimetry and energy transformations. Careful consideration of alternative symbols is required for eliminating existing ambiguities.

TRANSFORMATIONS IN OPEN SYSTEMS: INTERNAL AND EXTERNAL CHANGES

Open systems are treated most explicitly in nonequilibrium thermodynamics. Classical thermodynamics includes open systems, yet its main focus is on isolated and closed systems. The internal domain of any system is separated from the external domain (the surroundings) by a boundary. In theory, energy transformations outside the system can be ignored when describing the system. The surroundings are merely considered as a source or sink for quantities transferred across the system boundary. According to the transfer properties of the boundary, three types of thermodynamic systems are distinguished. (1) The boundaries of *isolated systems* are impermeable for all forms of energy and matter. Isolated systems do not interact with the surroundings. Strictly, therefore, internal changes of isolated systems cannot be observed from outside since any observation requires interaction. (2) The boundaries of *closed systems* are permeable for heat and work, but impermeable for matter. A limiting case is electrons which cross the system boundary when work is exchanged in the form of electric energy. The volume of a closed system may be variable. (3) The boundaries of *open systems* allow for the transfer of heat, work and matter. At present, IUPAC recommendations fall short of providing adequate guidelines for describing energy transformations in open systems.

Changes of isolated systems have exclusively internal origins, whereas changes of closed and open systems can be partitioned according to internal and external sources. Production and destruction of a quantity within the system are *internal* changes, whereas changes of heat, work and matter due to transfer across the system boundaries are labelled *external*. (External) transfer is thus contrasted with (internal) production or destruction. Accordingly, the system's entropy balance is (ref. 2),

$$\frac{dS}{dt} = \frac{d_i S}{dt} + \frac{d_e S}{dt} \quad (1)$$

The *exact differential*, dS [$J \cdot K^{-1}$], is the entropy change of the system which is thus a function of state. It is suggested here that subscripts following the symbol of the operator (d_i and d_e) indicate inexact differentials. The second law of thermodynamics restricts the internal entropy production (Eq. 1) to positive values, $d_i S \geq 0$. The external entropy, $d_e S$, may have either sign. For instance, heat flow from the system to the environment contributes to a negative external entropy. The external entropy may be negative to the extent that it overcompensates for the internal entropy production, which results in a decrease in the entropy of the system at the expense of the environment (ref. 3).

Some internal terms are restricted to zero by various conservation laws which rule out the production or destruction of the respective quantity. An example is the first law of thermodynamics, $d_i U = 0$. Note that *internal* (subscript i), as opposed to *external* (subscript e), must be distinguished from "internal-energy", U , which contrasts with "Helmholtz energy", A . The conservation of mass, m , in non-relativistic systems implies that $d_i m = 0$. Similarly, in the absence of nuclear reactions the amount of any atom is conserved, e.g. for carbon $d_i n_C = 0$. This is different for substances or chemical species which are produced or consumed in reactions, r (ref. 4). The change of entity i in an open system, dn_i , is due to both the internal formation of i , $d_i n_i = d_r n_i$ (positive if i is formed as a product within the system), and the external transfer of substance or chemical species i , $d_e n_i$ (negative if i flows out of the system and appears as a product in the surroundings),

$$\frac{dn_i}{dt} = \frac{d_r n_i}{dt} + \frac{d_e n_i}{dt} \quad (2)$$

Again, the exact time derivative, dn_i/dt [$mol \cdot s^{-1}$], pertains to changes of the system (no subscript following the operator d). For simplification, differentials will be used instead of the explicit time derivatives, until flows, fluxes and power are introduced (Eqs. 16 ff.). The exchange of matter, $d_e n_i$, is zero in closed systems. Then the external entropy transfer, $d_e S$ (Eq. 1), equals the exchanged heat, $d_e Q$ [J], divided by the absolute temperature of the external heat sink, T_e [K], at the heat-conducting system boundary. If there are several heat reservoirs in contact with the system, T_e refers to the heat sink with the lowest temperature relevant for heat exchange. In open systems, the additional transfer term of the entropy of matter, $d_{mat} S$, must be accounted for,

$$d_e S = \frac{d_e Q}{T_e} + d_{\text{mat}} S \quad (3)$$

The transfer of matter is associated with an exchange of enthalpy, $d_{\text{mat}} H$, Gibbs energy, $d_{\text{mat}} G$, and entropy, $d_{\text{mat}} S$, obtained by multiplying the amount of transferred substance by the partial molar quantity $H_{i,e}$, $G_{i,e}$ ($=\mu_{i,e}$) [$\text{J}\cdot\text{mol}^{-1}$] (Table 5A), and $S_{i,e}$ [$\text{J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$], respectively, at the conditions of the system boundary (at temperature, pressure and concentration of the external reservoir, indicated by subscript e), and relative to a chosen reference state (ref. 5,6).

$$d_{\text{mat}} H = \sum_i d_e n_i H_{i,e} \quad (4.1)$$

$$d_{\text{mat}} G = \sum_i d_e n_i G_{i,e} \quad (4.2)$$

$$d_{\text{mat}} S = \sum_i d_e n_i S_{i,e} \quad (4.3)$$

If there are various external reservoirs at different thermodynamic states, e.g. for the inflow and outflow of the system, then the matter-energies and entropies must be calculated separately for the conditions at each exchange site (ref. 6). An exception is the temperature of the heat sink (see below). In closed systems, the external matter-energy changes indicated by subscript mat (Eq. 4) are zero. External flows are quantified for the book keeping according to Eqs.(1) and (2). All external flows are zero in isolated systems. External heat flow, $d_e Q/dt$, and external power, $d_e W/dt$ (work per time [$\text{J}\cdot\text{s}^{-1}=\text{W}$]) are permitted in closed systems. None of these restrictions apply to the external flows in open systems. Note that external changes per time are *external flows* across the system boundaries which are formally reversible. Their irreversible facet is accounted for internally. Within the system boundaries, irreversible *internal flows* of heat and matter along gradients contribute to the internal entropy production, $d_i S$. In contrast, the external entropy change (Eq. 3) is not an expression of irreversibility or global entropy production (ref. 2).

The enthalpy change of the system, dH [J], is,

$$dH = d_i H + d_e H \quad (5)$$

At constant pressure, the distinction between enthalpy and internal-energy of a system is due to external pressure-volume work (Note a). Therefore, the internal terms are equal, $d_i U = d_i H$ ($p=\text{const.}$). The first law of thermodynamics states that internally internal-energy can be neither produced nor destroyed. This applies to conservation of enthalpy at constant pressure (ref. 6),

$$d_i H = 0; \quad p=\text{constant} \quad (6)$$

The system's enthalpy change, equal to the external enthalpy change, $dH = d_e H$, is the sum of the exchange of heat, work (excluding pressure-volume work; Note a), and matter (open system),

Note a: Available external work at constant pressure, $d_e W$, is here distinguished from external total work, $d_{\text{et}} W$. Defining pressure-volume work, $d_V W$, at constant pressure,

$$d_V W = -p dV$$

and hence enthalpy,

$$dH = dU - d_V W; \quad dp=0$$

the available work is,

$$d_e W = d_{\text{et}} W - d_V W$$

$$dH = (d_e Q + d_e W) + d_{\text{mat}} H \quad (7)$$

The first two *external* terms (in parentheses) on the right hand side of Eq.(7) are matched by the *internal* transformation of enthalpy, $d_i H$,

$$d_i H = d_e Q + d_e W \quad (8)$$

The type of energy transformation must be defined in each case. A partial transformation, tr (Eq. 20), is distinguished from the total of all simultaneous energy transformations, t (Eq. 8). If, for example, a chemical reaction, r , is the only energy transformation in an isothermal system, then $d_i H/dt = d_r H/dt$ is the reaction enthalpy per unit of time [$\text{J}\cdot\text{s}^{-1}=\text{W}$]. This is the case in an open system at steady state ($dH = 0$). The external flows of heat and work (Eq. 8) are quantitatively related to the chemical reaction enthalpy, and $d_i H = -d_{\text{mat}} H$ (Eq. 7). Temperature, pressure and concentrations are constant. Their changes would involve additional transformations summarized in $d_i H$. For example, the change of concentration of a compound is a chemical energy transformation between thermodynamic states, which may be included as a reaction, r , in general. The corresponding enthalpy change is negligible in many biological cases. Generally, by summing up all partial enthalpy transformations, Eq.(8) accounts for enthalpy transformations in any type of system.

Note the important difference between $d_i H$ (Eq. 8) and $d_r H$ (Eq. 6). Their equality is restricted to isolated systems. Transformations occur from one form of enthalpy to another, but formation and destruction are excluded. In closed and open systems, the transformed enthalpy may be exchanged as heat and work, hence the close association between external transfer and internal transformations at steady state ($dH = 0$; $d_i H = -d_{\text{mat}} H$). Closed systems are characterized by the equality $dH = d_i H$.

Contrary to Eq.(5), the system's change of Gibbs energy, dG (at constant p ; Gibbs energies, G , are substituted at constant volume by Helmholtz energies, A) contains a non-zero internal term for all irreversible energy transformations in the system (compare Eq. 1). This internal term is the *dissipated energy*, $d_i D$, integrated over all energy transformations (ref. 2),

$$dG = d_i D + d_e G \quad (9)$$

Conservation of Gibbs energy applies only to processes at equilibrium, when the potential for work is not destroyed (nor generated), $d_i D = 0$. This restriction to equilibrium or 100% efficiency is in contrast to the entirely general case of conservation of internal-energy (enthalpy; Eq. 6). The second law of thermodynamics restricts the dissipated energy to negative values, $d_i D \leq 0$. The sign is reversed when considering internal entropy production (Eq. 1; Note b),

$$-T_e d_i S = d_i D \quad (10)$$

The potential for work is lost fully or to some degree upon the advancement of irreversible or dissipative energy transformations. The irreversible destruction is expressed by a negative value of the dissipated energy, $d_i D < 0$. The external (reversible) Gibbs energy changes, $d_e G$, are accounted for by the transfer of work and matter (but not of isothermal heat; contrast with Eq. 7). Therefore, at constant pressure (Note a),

$$dG = (d_i D + d_e W) + d_{\text{mat}} G \quad (11)$$

External work, $d_e W$, is exchanged reversibly. In addition, the transfer term, $d_{\text{mat}} G$, does not refer to the irreversibility of the process. Therefore, irreversibility is exclusively an internal feature,

Note b: The *dissipated energy*, $d_i D$, emphasizes the importance of the relation to the entropy production, $d_i D = -T_e d_i S$. Confusion is avoided between $d_i D$ and $d_i G$ (Eq. 12), whereas the symbol $d_i G$ (for $d_i D$) places insufficient emphasis on this fundamental distinction. Introduction of the symbol $d_i D$ (versus $d_i G$) is justified as much as retaining the symbol $d_e W$ which in fact is a form of external Gibbs energy, too (Eqs. 11 and 12).

expressed by d_iD . The first two terms (in parentheses) on the right hand side of Eq.(11) are the irreversible (internal) and reversible (external) aspect of total Gibbs energy transformation, d_tG (Note b; contrast with the exclusively external terms on the right hand side of Eq. 8),

$$d_tG = d_iD + d_eW \quad (12)$$

Similarly, combining the corresponding internal and external terms (Eqs. 1, 3 and 10) yields, in analogy to Eqs.(8) and (12) the total *bound energy* transformation, $d_tB = T_e d_tS$ (ref. 5,7),

$$d_tB = -d_iD + d_eQ \quad (13)$$

The term 'bound energy' expresses the fact that this energy is not available or free for work but is (under isothermal conditions) bound to be exchanged as heat even at 100% efficiency of energy conversion, when $d_iD = 0$ (Eq. 13). Combining Eqs.(8, 12 and 13) yields the fundamental equation for energy transformations, which is also applicable for each partial transformation, tr,

$$d_{tr}H = d_{tr}G + d_{tr}B \quad (14)$$

HEAT AND WORK IN OPEN SYSTEMS

Gibbs energy transformation (Eq. 12) represents the distribution between internally dissipated energy and externally conserved energy (work). If work is performed *by* the system on the surroundings, $d_eW < 0$, then the *external efficiency* is, $\epsilon = d_eW/d_{in}G$. The input into the coupled process, $d_{in}G < 0$, is the exergonic Gibbs energy of an internal transformation. The second law restricts the efficiency to $\epsilon \leq 1$. Energy transformation may involve work performed *on* the system, $d_eW > 0$, in which case the external efficiency is, $\epsilon = d_{out}G/d_eW$. Then, the output of the coupled process, $d_{out}G > 0$, is the endergonic Gibbs energy of an internal transformation. The concept of external efficiency is extended to *internal efficiency*, considering energy conservation in the transformation from an internal input to a coupled internal output process. Therefore, the internal efficiency is (ref. 8), $\epsilon = -d_{out}G/d_{in}G$ (Table A6). The restriction, $\epsilon \leq 1$, due to the second law applies to all these internal and external efficiencies.

In contrast to Gibbs energy transformation (Eq. 12), enthalpy transformation (Eq. 8) expresses the partitioning between external heat and work. In some texts on nonequilibrium thermodynamics, heat and work have lost their well defined meaning. The pertinent complication does not arise from irreversibility but merely from a consideration of open systems. A confusing variety of definitions of heat and work stems from the treatise of open systems on the basis of Eq.(7), when the external term due to the transfer of matter, $d_{mat}H$, is omitted and arbitrarily split between 'heat' and 'work' (ref. 9,10). Then the term 'heat' is used for (1) external entropy, $T_e d_eS = d_eQ + T_e d_{mat}S$ (Eq. 3) and 'work' is $d_eW + d_{mat}G$ (compare Eq. 11); or (2) the sum $d_eQ + d_{mat}H$ in which case work is d_eW ; or (3) 'pure heat' (ref. 9), d_eQ , in which case 'work' contains the enthalpy (!) of transferred matter, $d_eW + d_{mat}H$ (ref. 10). From the perspective of experimental calorimetry, it appears problematic to accept these definitions of 'various heats', which lead to terminological if not conceptual confusion. Moreover, the fundamental restriction, $\epsilon \leq 1$, of the second law would brake down for the external efficiency when defining 'work' in terms of cases (1) or (3). An unambiguous and general definition of external heat and work must account for the fundamental distinction between the energy changes in transformations, d_tH and d_tG (which contain *differences of potentials*, see below), and the *relative* energy changes in the transfer or matter, $d_{mat}H$ and $d_{mat}G$, which lack an absolute value since they contain *potentials* which depend on the arbitrary choice of reference states (Eq. 4).

Heat exchange is restricted to conduction and radiation in closed systems. In addition, convection of heat is combined with the exchange of matter in open systems. How is it then possible in Eq.(7) to separate d_eQ and $d_{mat}H$? Conceptual clarification may be aided by considering calorimetric experiments. In *example 1*, let a chemical reaction advance as the only energy transformation in a

perfusion heat-flow calorimeter (ref. 11). Heat is exchanged through the heat-conducting system boundary and by convection through the inflow and outflow of matter. At steady state and in the absence of external work, $d_e W = 0$, the isothermally exchanged heat is a result of the internal transformation of chemical energy, $d_r H$, irrespective of the relative contributions to $d_e Q$ of conduction, radiation and convection ($d_r H = d_e Q$; Eq. 8). Convective heat exchange is, therefore, not included in $d_{\text{mat}} H$ but is fully accounted for in the heat term, $d_e Q$ (Eq. 7). This is straightforward under ideal isothermal conditions when the inflow and outflow of matter actually occur at the temperature of the external heat sink, T_e . Consider, that different temperatures, T_j , are involved in heat transfer. Let the temperature differences, $T_j - T_e$, be small such that the molar heat capacities of substances i at constant pressure, $C_{i,p}$ [$\text{J}\cdot\text{K}^{-1}\cdot\text{mol}^{-1}$], do not change over the respective temperature range. In this case, convective heat transfer associated with the exchange of matter, $d_{\text{mat}} Q$, is,

$$d_{\text{mat}} Q = \sum_i \sum_j d n_{ij} C_{i,p} (T_j - T_e) \quad (15)$$

The summation is over all exchanged substances i and over temperatures j at which substance i is transferred at temperature T_j . The convective heat flow (Eq. 15) is added to the conductive and radiative heat flow to obtain the total external heat flow, $d_e Q/dt$. As a consequence, the matter-energy change is correctly calculated for all substances with reference to the same T_e (Eq. 4), even if the actual temperature at the specified exchange site is T_j . In fact, this is achieved directly when calibrating a heat flow calorimeter in perfusion mode (example 2).

For ensuring defined conditions of temperature at steady state, it is of primary importance that an energy transformation is associated with a well defined heat exchange, whereas the mode of heat exchange is of secondary relevance. Experimental consideration of the mode of heat exchange is critical, whence constant convective and conductive proportions of external heat flow must be established during calorimetric measurements and calibration. In *example 2*, let external electrical power be supplied for calibration of the perfusion heat-flow calorimeter, and the corresponding external electric energy be dissipated internally across a calibration resistor. The efficiency of the transformation of externally supplied work is zero, $d_{\text{out}} G = 0$, and also no other internal transformations occur, $d_t G = 0$ (therefore, $d_i D = -d_e W$; Eq. 12). Like the total Gibbs energy transformation, the total (net) enthalpy transformation is zero, $d_t H = 0$, since any internal electric charging is simultaneously compensated by internal electric discharging at steady state (therefore, $d_e Q = -d_e W$; Eq. 8). In the absence of any other internal processes ($d_t B = 0$), the dissipated energy equals the heat exchange, $d_e Q = d_i D$ (Eq. 13), when the externally supplied electric work is internally dissipated. Again, this yields an unambiguous definition (calibration) of heat exchange on the basis of electric energy dissipation, which applies unequivocally to closed and open calorimetric systems.

In open systems, the actual values of the external matter-energy changes (Eq. 4) depend on the largely arbitrary choice of the reference state for the partial molar quantities, $H_{i,e}$, $G_{i,e}$ and $S_{i,e}$. Values for these quantities can only be given by assuming a *theoretical* transformation from the actual state of i to a reference state. The reference states of formation or combustion are conventional choices (ref. 5,6). In contrast, the *actual* transformation energy changes are independent of the choice of reference states. Reference values may be used for computation, but these cancel out when calculating the difference between the initial and final states which define the transformation. Therefore, an unequivocal definition of heat and work can only be based on the transformation energy changes. Separation of the transformation energy changes from the external matter-energy changes (Eqs. 7, 11) provides unique theoretical advantages and offers a practical and uniform definition of heat and work in closed and open systems.

To emphasize this point, approaches based on Eqs.(8 and 12-14) are labelled as process or *transformation analysis* (ref. 7), in comparison to a complete *system analysis* (Eqs. 1, 7, and 11). The two approaches converge in closed systems, where the energy changes of the system are identical to the transformation energy changes ($dH = d_t H$, Eqs. 7-8; $dG = d_t G$, Eqs. 11-12; $T_e dS = d_t B$, Eq. 13).

Consider (*example 3*) the reaction of *example 2* now to proceed in a thermally isolated, adiabatic system ($d_e Q = 0$) in the absence of external work ($d_e W = 0$). Under these conditions, $d_t H = 0$ (Eq. 8). This requires that the (partial) reaction enthalpy, $d_r H$, is compensated by the additional (partial) thermal enthalpy, $d_{th} H = -d_r H$, resulting in a total enthalpy transformation of zero, $d_t H = d_r H + d_{th} H$. The thermal enthalpy change is the product of the heat capacity, C_p [$J \cdot K^{-1}$], and the temperature change, $d_{th} H = C_p dT$. When heat-flow calorimeters are operated at non-steady states with respect to temperature (*example 4*), it is conventional to calculate a 'time correction' of the heat signal (ref. 11). This yields the total heat flow, $d_t Q/dt$ [$J \cdot s^{-1} = W$], including the external heat flow, $d_e Q/dt$, and the internal flow of thermal enthalpy, $d_{th} H/dt$,

$$\frac{d_t Q}{dt} = \frac{d_e Q}{dt} + C_p \frac{dT}{dt} \quad (16)$$

External heat flow, $d_e Q_j/dt$, at heat exchange sites with temperatures $T_j \neq T_e$, is different from external thermal power which is the transfer of thermal work per unit time, $d_{th} W/dt$,

$$\frac{d_{th} W}{dt} = \sum_j \frac{d_e Q_j}{dt} \frac{T_j - T_e}{T_j} \quad (17)$$

Thermal work is not accounted for in $d_{mat} G$ (which must be calculated at T_e) even when convection of heat at T_j contributes to $d_e Q_j$ (see also Tables A4 and A6). A Carnot heat engine can conserve at maximum efficiency all the external thermal power input, $d_{th} W_{in}/dt$, and convert it into another form of external power output, $d_{et} W_{out}/dt$, including pressure-volume work (see Note a). At zero efficiency, the thermal work input is lost as dissipated energy, $d_i D$. Dissipation of thermal potential energy (exergy) in non-isothermal processes requires a distinction between the temperature of the external heat sink, T_e [K], and the temperature at which the j th energy transformation takes place, T_j . In this case the dimensionless temperature ratio, T_e/T_j , is used as an integration factor for obtaining the dissipated energy, $d_i D$, replacing the conventional integration factor $1/T_j$ which yields entropy (ref. 7). Gradients of temperature or chemical potential may exist across the actual wall (e.g. a glass wall or a biological membrane) of an enclosed experimental system. Then the wall is considered as part of the system, or the wall is described as an additional (sub)system. The physical dimension of the selected boundaries must be negligible in relation to the macroscopic system, such that dissipation or irreversibility occurs either in a system or in the surroundings.

IRREVERSIBLE PROCESSES, INTERNAL FLOWS, AND ADVANCEMENT

The balance equation (Eq. 2) can be written for a range of extensive quantities X_i , which are converted in energy transformations, t , and externally exchanged, e ,

$$\frac{dX_i}{dt} = \frac{d_t X_i}{dt} + \frac{d_e X_i}{dt} \quad (18)$$

dX/dt is known as the 'rate of change of quantity X ' (ref. 1, p. 49, see also p. 58). $d_e X_i/dt$ is the external flow of quantity X_i in the thermodynamic state of entity i (ref. 6). A positive external flow explicitly takes into account that, as the quantity $d_e X_i$ is added to the system, it simultaneously disappears from the surroundings, both at the thermodynamic state pertaining to the system boundary (external reservoir). In Eq.(18), the subscript t indicates all transformations involving quantity X_i . Similarly, every partial energy transformation, tr , can be defined as a balance of production and destruction (chemical reaction including phase changes), or local appearance and removal (transport between locations of different potential). In a spontaneous *transformation flow*, the quantity $d_{tr} X_2$ is a product formed at potential 2 internally, from a source which disappears simultaneously within the system, $-d_{tr} X_1$, at potential 1. In the simplest case, a transformation flow can be quantified by reference to either the homogenous side 1 or 2 of the process,

$$0 = v_1 X_1 + v_2 X_2 \quad (19)$$

In discontinuous systems, e.g. diffusion through a membrane from a homogenous side 1 to the homogenous side 2, the balance equation (Eq. 19) holds only at steady state, when no accumulation occurs in the diffusion path. This is comparable to the application of the balance equation (Eq. 19) to a chemical reaction, $0 = -1 A + 1 B$, which holds only if no intermediary complexes (e.g. AB) accumulate. ν_i is known as the stoichiometric number of entity i in chemical reactions (ref. 1). For diffusion of a substance through a membrane, $\nu_1 = -1$ would indicate that the direction of positive flow is defined from side 1 (left) to side 2 (right), the reverse being expressed by a negative flow. For chemical reactions, the extent of reaction or advancement, $d_r \xi = d_r n_i \nu_i^{-1}$ (ref. 1), is side-independent, that is, the same result is obtained when observing either reactant A or product B for calculating the reaction flow. Therefore, the side-independent expression of any defined type of transformation, tr , requires a generalization of the concept of *advancement*, $d_{tr} \xi$ (Table 1). This generalization is an extension of the advancement of reactions (ref. 4) and the advancement of the transport of charged components, introduced by Prigogine (ref. 2). For any transformation, tr ,

$$d_{tr} \xi = \frac{d_{tr} X_i}{\nu_i} \quad (20)$$

The subscript in X_i relates either to an initial or final state of quantity X , the high-potential source, 1, or the low-potential sink, 2. The thermodynamic state for X_i (temperature, pressure, composition) is a function of time in non-steady state systems. In general, the advancement is $d_{tr} X_i$ normalized for ν_i (Eq. 20). ν_i is positive for products or for X_i written on the right hand side for a process indicated to proceed from left to right in a diagram. If the product (on the right hand side, with positive ν) is actually removed, then the advancement (Eq. 20) has a negative sign indicating that the flow is directed backwards, from right to left. The type of energy transformation is specified by an index for the transformation (t, total of simultaneous transformations; tr, partial transformation in general; r, reaction; d, diffusion; el and th, electrical and thermal conduction), and by the property X involved in the process (number of particles, mass m , substance B, charge Q_{el} , heat Q). Accordingly, *transformant* was suggested as the name of quantity X , related to the advancement of transformations (John Schellmann, IUPAC workshop 1992).

The *balance approach to transformations* between internal sources and internal products, has important advantages compared to the mere statement of an internal formation or destruction. The transformation contains explicitly the link between the product and the internal source. This is particularly important for an intuitive understanding of internal entropy production or energy dissipation. Entropy originates from the dissipation of an ordered state. In other words, the potential for doing work can only be destroyed (negative value of the dissipated energy, $d_i D$), if a potential for an energy transformation was there in the first place.

De Donder (ref. 4) introduced the concept of advancement for chemical reactions, $d_{tr} \xi = d_r \xi$ [mol]. The stoichiometry must be specified (in the form of Eq. 19 for a reaction $0 = -A + B$). A stoichiometric form may be chosen such that the stoichiometric number of substance B is unity (1 or -1), for example in a reaction $0 = -B + 2C$. Then the advancement $d_r \xi_B$ is expressed in equivalents of the amount of B. The stoichiometric number of i takes account of the fact that internally the entity i is either removed in the process ($\nu_i < 0$ for reactants) or produced ($\nu_i > 0$ for products).

Substituting the advancement (Eq. 20) into Eq.(18) for quantity X_i , in systems involving only one transformation tr , yields the balance equation in terms of internal and external flows,

$$\frac{dX_i}{dt} = \frac{d_{tr} \xi}{dt} \nu_i + \frac{d_e X_i}{dt} \quad (21)$$

$d_e X_i / dt$ is a reversible external or *transfer flow* of quantity X_i in a thermodynamic state pertinent to the system boundary. Since transfer does not include transformation, the balance (Eq. 21) must be written separately for any educt and product of a transformation. The transfer flows account for the exchange of entropy and are formally separated from internal entropy production (Eq. 1). Irreversibility is expressed by internal entropy production, exclusively associated with internal flows (ref. 2).

TABLE 1: Advancement of energy transformations, $d_{tr}\xi$, related to the flows, I_{tr} , by the relation, $I_{tr} = d_{tr}\xi/dt$.

<i>name</i>	<i>symbol</i>	<i>relation</i>	<i>SI unit</i>
Advancement	$d_{tr}\xi$	$= d_{tr}X_i v_i^{-1}$	x
Electric advancement	$d_{el}\xi$	$= d_{el}Q_i v_i^{-1}$	C
Thermal advancement	$d_{th}\xi$	$= d_{th}Q_i v_i^{-1}$	J
Diffusion advancement	$d_d\xi_B$	$= d_d n_i v_i^{-1}$	mol
Reaction advancement	$d_r\xi_B$	$= d_r n_i v_i^{-1}$	mol

$d_{el}Q_i$ ($d_{th}Q_i$) are the changes in electric charge (heat) at the locations of high or low electric potential (temperature) within the system.

As discussed in examples 1 and 4, $d_e Q_j/dt$ describes isothermal (locally reversible) external heat flow across the system boundary at the boundary temperature T_j , e.g. at the high- (1) and low-temperature (e) heat exchange site, $d_e Q_1$ and $d_e Q_e$. At steady state, the temperatures in the various compartments of the system are constant, whence the internal and external changes compensate each other during any time interval, $0 = d_{th}Q_1 + d_e Q_1$ (ref. 7). Heat flow along a temperature gradient is an (internal) thermal transformation flow, $d_{th}\xi/dt$ (Table 1). This internal heat flow relates to the amount of heat leaving the internal high temperature region, $d_{th}Q_1 < 0$ ($v_1 = -1$; Eq. 19) which equals, at zero efficiency, the amount of heat added to the internal low temperature region, $d_{th}Q_e > 0$ ($v_e = +1$; Eq. 19). In the case of coupling of the thermal transformation to an output process, there is conservation of thermal energy (see also Table A6), and less heat is added to the low temperature region ($v_e < 1$). Therefore, internal or external work, such as lifting a weight or charging a capacitor within the system or outside, must be considered as an additional quantity in the balance equation. In line with the second law of thermodynamics, isothermal heat flow cannot be transformed into power output, and the thermal power is zero in isothermal heat flow when $T_e - T_1 = 0$ (Eq. 22). In the absence of other processes, the internal **thermal power**, P_{th} , yields the entropy production with the temperature of the external heat sink, T_e , chosen as a reference (ref. 7; compare Eq. 17 for external thermal power, $d_{th}W/dt$),

$$-T_e \frac{d_i S}{dt} = P_{th} = \frac{d_{th}\xi}{dt} \frac{T_e - T_1}{T_1} \quad (22)$$

FORCES: A CLOSE LOOK AT EXTENSIVE AND INTENSIVE QUANTITIES

1. **Extensive quantities** refer to a defined system. They change in proportion to the size of a homogenous system (Table A1). There are two fundamentally different ways to change extensive quantities into those which are independent of system size (and which are traditionally combined as intensive properties). These yield either size-specific or process-intensive quantities.

2. **Specific quantities** are obtained when the extensive quantity is divided by system size (Table A2). For example, division by the mass or volume of the system yields *mass-specific* or *volume-specific* quantities. A molar quantity (divided by amount of substance, ref. 1, p. 7) may be size-specific. For example, in a system containing a given amount of organic carbon, the extensive quantity Gibbs energy, G [J], divided by the amount of organic carbon yields the specific molar quantity G_{nC} [J.C-mol⁻¹] (Table A2). The term *specific* is used here in the general sense of 'size-specific', extending the IUPAC definition where 'specific' is restricted to 'mass-specific' (ref. 1, p. 7).

TABLE 2. Ergodynamic forces, F_{tr} , related to the advancement of energy transformations, tr, by the relation, $F_{tr} = \partial G / \partial_{tr} \xi$, in discontinuous systems. These scalar forces are potential differences.

name	symbol	relation	SI unit
Ergodynamic force	F_{tr}		$J \cdot x^{-1}$
Electric force	F_{el}	$= V_2 - V_1$	$J \cdot C^{-1} = V$
Thermal force	F_{th}	$= (T_e - T_1) / T_1$	1
Force of diffusion of B	$F_{B,d}$	$= \mu_{B2} - \mu_{B1}$	$J \cdot mol^{-1}$
Gibbs force of reaction	$F_{B,r}$	$= \sum \nu_i \mu_i$	$J \cdot mol^{-1}$

V_1 and V_2 are the potentials of the electrodes at an infinitely small electric flow, shown on the left (1)- and right (2)-hand sides in a circuit diagram (ref. 1).

μ_{B1} and μ_{B2} are the chemical potentials of substance B in compartments 1 (left) and 2 (right), where positive flow of B is defined in the direction from 1 to 2.

With $\nu_{B1} = -1$ and $\nu_{B2} = +1$, the equation

$$F_B = \sum \nu_i \mu_i$$

is generalized for reaction and diffusion.

3. *Intensive quantities* are partial derivatives of an extensive quantity by the advancement, $d_{tr} \xi$, of an energy transformation. The term *intensive* is used here in the restricted sense of 'process-intensive'. Particularly important intensive quantities are the partial derivatives of Gibbs energy (or Helmholtz energy) by the advancement of a specified energy transformation, tr. These are the *ergodynamic forces* (Table 2; Note c),

$$F_{tr} = \frac{\partial G}{\partial_{tr} \xi} \quad (23)$$

Some ergodynamic forces are known in classical thermodynamics as 'electromotive forces' or 'potential differences' (Table 2): electric potential difference, ΔV , electrochemical potential difference, $\Delta \bar{\mu}$, chemical potential difference, $\Delta \mu$. The ergodynamic forces must be distinguished from the fundamental forces of physics. The electric potential (ref. 1), $V = dW/dQ$ [$J \cdot C^{-1}$], and chemical potential (ref. 1), $\mu_B = (\partial G / \partial n_B)_{T, p, n_{i \neq B}}$ [$J \cdot mol^{-1}$], are partial derivatives formally similar to Eq.(23). However, ergodynamic forces are Gibbs energy changes per advancement from one actual to another actual state in the system (Eq. 23). In contrast, potentials are Gibbs energy changes per advancement from a purely theoretical reference state to the actual state in the system. The reference states have zero energies by definition. The customary zero potential for charge is defined at infinite distance of the charged particles, and zero chemical potentials are frequently defined in terms of the reference state of formation (the elements at their standard states).

For reactions, the partial derivative of the extensive quantity G by the advancement yields the reaction Gibbs energy per mol advancement, $\Delta_r G_B$ [$J \cdot mol^{-1}$] (where reaction r is implicitly defined such that the stoichiometric number of substance B is unity). The appropriate term for this intensive

Note c: X indicates a generalized force (*intensive* quantity) in nonequilibrium thermodynamics (ref. 2), but a general *extensive* quantity in IUPAC (ref. 1, p. 43). To avoid ambiguity, the generalized forces, X_{tr} , associated with entropy production, are substituted by ergodynamic forces relating to power, $F_{tr} = -X_{tr} T$ (Table 2). 'Ergodynamics' is a short and accurate term for nonequilibrium thermodynamics. The important distinction is emphasized between *thermodynamics* related fundamentally to heat and *ergodynamics* related strictly to potential energy changes (erg, work; Eq. 23).

quantity is *Gibbs force* of reaction, $F_{B,r}$ (Table 2; ref. 5),

$$F_{B,r} = \left(\frac{\partial G}{\partial \xi_B} \right)_{T,p} = \sum_i \nu_i \mu_i ; \quad |\nu_B| = 1 \quad (24)$$

The symbol $\Delta_r G_B$, suggested as an alternative for $F_{B,r}$ more in line with previous conventions, expresses the fact that the Gibbs force is a difference of *stoichiometric potentials*, $\nu_i \mu_i$ [$\text{J}\cdot\text{mol}^{-1}$], of products and substrates with positive and negative stoichiometric numbers, ν_i , respectively. The IUPAC symbol, based on an explicitly defined reaction stoichiometry, is $\Delta_r G_m$ (ref. 1). The symbol ΔG for the partial molar quantity lead to criticism owing to the fact that it indicates a difference rather than a differential quotient (ref. 5). Both the differential quotient and the difference (summation) are justified (Eq. 24) (see also ref. 10). However, the discrimination between the Gibbs energy of reaction, $\Delta_r G$ [J] (extensive), or Gibbs energy of reaction per volume, $\Delta_r G_V$ [$\text{J}\cdot\text{m}^{-3}$] (specific), and Gibbs force, $F_{B,r}$ [$\text{J}\cdot\text{mol}^{-1}$] (intensive), warrants basically different symbols. Importantly, $F_{B,r}$ ($\Delta_r G_B$) is *not* an *energy* of reaction but the driving *force* of the reaction (ref. 5,8).

It is common practice to use a single set of symbols for the extensive (Table A1) and the corresponding molar quantities. In this manner, the number of symbols is reduced. In view of the fundamental distinction between size-specific molar quantities and process-intensive molar quantities, such a simplification is acceptable only when restricted to the specific quantities. Based on the same argument and selecting a different measure of system size (mass or volume instead of amount of substance), it may be practical to use the symbols of the extensive quantities for their specific (mass-specific or volume-specific) counterparts (ref. 5). When it is necessary to discern the extensive and corresponding specific quantities explicitly, application of more complex symbols cannot be avoided (Table A2). In any case, the process-intensive quantities should be clearly distinguished from the extensive and specific properties.

FLOWS AND FORCES RELATED TO ENTROPY AND POWER

The names 'irreversible' or 'nonequilibrium' thermodynamics suggest a contrasting restriction of classical thermodynamics to processes at equilibrium. In this regard they are misnomers. The emphasis on energy transformations related to work (*ergon*) and power as opposed to heat (*therme*) is expressed by the term *ergodynamics* (ref. 5,7). Ergodynamic consideration of flows and forces is based on energy transformations in the sense of the second law.

Power is work per unit time (external) or Gibbs energy per unit time (internal). Internal electric power, P_{el} [$\text{W}=\text{J}\cdot\text{s}^{-1}$], is,

$$P_{el} = \frac{d_{el}G}{dt} = I_{el} F_{el} = \frac{d_{el}\xi}{dt} \frac{\partial G}{\partial_{el}\xi} \quad (25)$$

$d_{el}G$ is the electric part of the Gibbs energy transformation (Table A3). For any energy transformation, the partial Gibbs energy change, $d_{tr}G$ [J], is the product of the advancement, $d_{tr}\xi$, and the corresponding ergodynamic force, F_{tr} ,

$$d_{tr}G = d_{tr}\xi F_{tr} \quad (26)$$

Adopting the symbol I for electric current, the flow of a transformation is I_{tr} (Note a),

$$I_{tr} = \frac{d_{tr}\xi}{dt} \quad (27)$$

Note a: *Flow, flux* and *rate* are used synonymously in nonequilibrium thermodynamics, with symbol J . This contradicts the IUPAC restriction of J for vector fluxes (ref. 1, p. 58). Distinct from flux, J (size-specific), the symbol I is adopted from the IUPAC symbol for electric current for generalized flows (extensive).

Based on ergodynamic forces as partial derivatives of Gibbs energy (Eq. 23), the sum of flow-force products yields the power of total energy transformation in the system, P_t [W],

$$P_t = \sum_j I_j F_j \quad (28)$$

A spontaneous, exergonic process within the system is associated with a dissipation of Gibbs energy. Hence a negative power, $P_{tr} < 0$ [W], is associated with a positive spontaneous flow and a negative ergodynamic force (Eq. 28). The sign of the flow depends on the definition of the forward (positive) direction of the process. For a chemical reaction, $A \rightarrow B$, the sign of a given flow is automatically reversed when the reaction is written as $B \rightarrow A$, since the signs of the stoichiometric numbers are changed (Eqs. 20 and 27). Simultaneously, the sign of the Gibbs force changes (Eq. 24). As a result, the flow-force product (Eq. 28) is independent of the arbitrary choice of direction.

In nonequilibrium thermodynamics, the product of flows and forces yields the rate of entropy production, $d_i S/dt$ (ref. 2; see Note c). From Eq.(12) it is shown that this definition of entropy production is restricted to systems on which no external work is performed, $d_e W = 0$. According to Eq.(12), power is split into an internal irreversible and an external reversible part,

$$P_t = P_i + P_e \quad (29)$$

$P_i = d_i D/dt \leq 0$ is related to the entropy production (Eq. 10). External power input, $P_e > 0$, can drive a flow against its intrinsic force in an endergonic process which may lead to a positive balance of Gibbs energy transformation, $P_t \leq P_e$.

FLOWS AND FLUXES

The forces of nonequilibrium thermodynamics represent "affinities, gradients of temperature and of chemical potentials, etc." (ref. 13). Including gradients or vector forces in Eq.(28) generates a dimensional problem (flow times force must yield power; unit [W]). This problem is solved by distinguishing carefully flows, I , and fluxes, J , the latter leading to a local formulation of power dissipation per unit of volume, known as the dissipation function (ref. 9).

The IUPAC terminology for electricity is very detailed and may serve as a guideline. Flow (current), I , and flux (current density), J , are clearly distinguished (Table A3). In the IUPAC terminology for transport properties, flux is defined as flow per unit of area (Note e),

$$J_{tr} = \frac{d_{tr} \xi}{dt} \frac{1}{A} \quad (30)$$

The symbols $d_{tr} Y$ and $d_{tr} Y$ are suggested for the specific advancement per unit area and volume, respectively (Table A2). The distinction between extensive and size-specific expressions of advancement is as important as the separation between flow and flux. This discrimination is extended to a distinction of scalar flows and fluxes. In a scalar transformation, flux is flow divided by volume,

$$J_{tr} = \frac{d_{tr} \xi}{dt} \frac{1}{V} \quad (31)$$

Note e: The difference between *flow* and *flow rate* (ref. 1) is unclear. Considering the important distinction between population growth, $I_g = dN/dt$ (extensive), and population growth rate, $r_g = (dN N^{-1})/dt$ (size-specific), it is useful to restrict *rate*, r_{tr} , to quantities with the dimension per time [s^{-1}]. For vector flux the symbol \mathbf{J} in bold rather than J is chosen in line with the symbol for vector electric flux (current density; ref. 1). Then J indicates the scalar flux (per unit volume).

Each conjugated flux-force pair is strictly interrelated: Their product yields volume-specific power,

$$P_{V,t} = \sum_j J_j F_j + \sum_k J_k F_k \quad (32)$$

where the summation is over all internal processes (j , scalar; k , vector). As with the extensive treatment of power (Eq. 29), the volume-specific power, $P_{V,t}$, is split into the internal irreversible dissipation function, $\Phi = P_{V,i} = d_i D_V/dt$, and the external reversible power per volume,

$$P_{V,t} = \Phi + P_{V,e} \quad (33)$$

Equality of power dissipation (the dissipation function; Note f) and the sum of flux-force products applies exclusively to processes with zero external power, $P_{V,e} = 0$. Consider a battery during the process of charging as a closed system. External electric power, $P_{V,e} > 0$, drives the internal electrochemical flux, $J_r > 0$, against an endergonic Gibbs force, $F_r > 0$. Then the product of flux and force yields a positive internal power, $P_{V,r} > 0$. Part of the external power input is dissipated within the battery, $\Phi < 0$ (Note f). The dissipative process of charging a battery, therefore, yields a positive total power due to transformation of work and Gibbs energy (Eq. 33). If the system boundary is redrawn as to enclose the electric power input as an internal partial process, then $P_{V,t} = \Phi$ (Eq. 33), in which case the dissipation function equals the sum of flux-force products (Eq. 32), $\Phi = J_{el} F_{el} + J_r F_r < 0$.

METABOLIC FLUX, HEAT FLUX AND POWER

Reaction power per volume, $P_{V,r}$ [$\text{W}\cdot\text{m}^{-3}$], is the product of chemical flux (Eq. 31), $J_B = d_r \xi_B/dt \text{ V}^{-1}$ [$\text{mol}\cdot\text{s}^{-1}\cdot\text{m}^{-3}$] and Gibbs force, F_B [$\text{J}\cdot\text{mol}^{-1}$],

$$P_{V,r} = J_B F_B \quad (34)$$

Chemical flux can be measured calorimetrically when the enthalpy change of the reaction system is only due to heat changes without any transfer of work. See Eq.(16) for the definition of heat flow, I_Q . Under these conditions, scalar heat flux, J_Q [$\text{W}\cdot\text{m}^{-3}$], and enthalpy flux due to the reaction, $J_{H,r}$, are equal. They are related to the chemical flux, expressed in terms of volume-specific advancement of reaction, by the molar reaction enthalpy, $\Delta_r H_B$ [$\text{J}\cdot\text{mol}^{-1}$] (Table A5),

$$J_Q = J_{H,r} = J_B \Delta_r H_B \quad (35)$$

Substituting Eq.(35) into (34), chemical reaction power is calculated from the enthalpy (or heat) flux on the basis of the Gibbs force normalized by the molar reaction enthalpy, at $P_{V,e} = 0$ (ref. 7),

$$P_{V,r} = J_Q \frac{F_B}{\Delta_r H_B} \quad (36)$$

Therefore, heat flow (heat flux) and (specific) power (ref. 14) must be clearly distinguished, analogous to the important contrast between reaction enthalpy and reaction Gibbs energy. The ratio $F_B/\Delta_r H_B$ is very different from unity in many important biochemical reactions. Metabolic *heat flux* in biological energy transformation is related to the chemical fluxes by the molar reaction enthalpies (Eq. 35), whereas specific metabolic *power* is related to the chemical fluxes by the *Gibbs forces* (Eq. 34). The distinction of heat flow and power is of fundamental importance when discussing the (ergodynamic) power efficiency of biological energy transformations (ref. 8; Table A6).

Note f: The dissipation function has a reversed sign when defined as the product of absolute temperature and volume-specific internal entropy production (ref. 9).

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REFERENCES

1. I. Mills, T. Cvitas, K. Homann, N. Kallay and K. Kuchitsu, *IUPAC. Quantities, Units and Symbols in Physical Chemistry*. Blackwell, Oxford (1988).
2. I. Prigogine, *Introduction to Thermodynamics of Irreversible Processes*, Interscience, New York, 3rd ed. (1967).
3. E. Schrödinger, *What is Life?* Cambridge University Press (1944).
4. Th. De Donder and P. Van Rysseberghe, *Affinity*. Stanford Univ. Press, Calif. (1936).
5. E. Gnaiger, *Thermochim. Acta* **172**, 31-52 (1990).
6. U. von Stockar, E. Gnaiger, L. Gustafsson, I. Marison, C. Larsson and P. Tissot, *Biochim. Biophys. Acta* (in press).
7. E. Gnaiger, *Thermochim. Acta* **151**, 23-34 (1989).
8. E. Gnaiger, In: *Surviving Hypoxia: Mechanisms of Control and Adaptation*. Eds. Hochachka P.W., Lutz P.L., Sick T., Rosenthal M., Van den Thillart G., CRC Press, Florida, pp. 77-109 (1993).
9. A. Katchalsky and P.F. Curran, *Nonequilibrium Thermodynamics in Biophysics*. Harvard Univ. Press, Cambridge, Mass. (1965).
10. H.V. Westerhoff and K. Van Dam. *Thermodynamics and control of free-energy transduction*. Elsevier, Amsterdam (1987).
11. J. Suurkuusk and I. Wadsö, *Chem. Scripta* **20**, 155-163 (1982).
12. G.R. Welch, *J. theor. Biol.* **114**, 433-446 (1985).
13. I. Prigogine, *From Being to Becoming. Time and Complexity in the Physical Sciences*. p. 85, Freeman, New York (1980).
14. I. Wadsö (ed.), *Pure & Appl. Chem.* **58**, 1405-1410 (1986).
15. D.H. Whiffen, ed. *IUPAC. Manual of Symbols and Terminology for Physicochemical Quantities and Units*, Pergamon Press, Oxford (1979).

APPENDIX: LIST OF SYMBOLS

IUPAC nomenclature and symbols are reported in the *notes* for comparison (ref. 1) with short explanations of the modifications made as far as necessary.

TABLE A1. Extensive quantities in energy transformation and the corresponding force. Extensive quantities relate to a system which must be defined in terms of its boundaries and size. The subscript tr indicates a partial internal transformation.

<i>name</i>	<i>symbol</i>	<i>relation</i>	<i>SI unit</i>	<i>notes</i>
1 power	P_{tr}	$= d_{tr}G/dt = I_{tr} F_{tr}$	$W=J\cdot s^{-1}$	1
2 partial Gibbs energy of transformation	$d_{tr}G$	$= d_{tr}\xi F_{tr}$	$J=kg\cdot m^2\cdot s^{-2}$	2
3 ergodynamic force	F_{tr}	$= \partial G/\partial_{tr}\xi$	$J\cdot x^{-1}$	3
4 flow	I_{tr}	$= d_{tr}\xi/dt$	$x\cdot s^{-1}$	4
5 advancement	$d_{tr}\xi$	$= d_{tr}X_i v_i^{-1}$	x	5
6 transformant	X_{tr}		x	6
7 conductance	R_{tr}^{-1}	$= -I_{tr} F_{tr}^{-1}$	$x^2\cdot W^{-1}$	7

- (1) IUPAC: (mechanical) power, $P = dW/dt$.
External power, $P_e = d_eW/dt$, is contrasted with P_{tr} , the power of an *internal energy transformation*, negative for exergonic and positive for endergonic transformations. When the total transformation, t , accounts for all simultaneous internal processes in the system and $d_eW = 0$, then P_t equals the dissipation function, $\Phi \leq 0$. The value of Φ is negative in the context of power dissipation (positive in the context of internal entropy production).
- (2) IUPAC: Gibbs energy, $G = H - TS$; Helmholtz energy, $A = U - TS$. An exergonic (spontaneous) change has a negative sign.
 $d_{tr}G$ is most useful at constant pressure, $d_{tr}A$ at constant volume.
- (3) IUPAC examples: (electricity) electromotive force, E ; electrical potential difference, U , ΔV , $\Delta\phi$; (electrochemistry) electromotive force, E ; electric potential difference, ΔV , E , U ; (chemistry) affinity of reaction, reaction Gibbs energy, $-A = \Delta_r G$.
Scalar forces in discontinuous systems are potential differences. These may be scalar products of two vectors, $F_{tr} = \int F_{tr} \cdot ds$. Examples for such 'motive forces' in discontinuous systems are the electromotive force, $F_{el} = \int (F/Q_{el}) \cdot ds$, and the difference of chemical potential of substance B across a membrane, $F_{B,m} = \Delta\mu_B = \int (\partial\mu_B/\partial s) \cdot ds$.
- (4) IUPAC examples: electric current, I ; mass flow rate, q_m ; heat flow rate, ϕ ; (chemical) rate of conversion, $\dot{\xi}$.
- (5) IUPAC example: (chemical) extent of reaction, advancement, ξ .
 x is the symbol for units which vary according to the nature of X in the transformation.
- (6) IUPAC: quantity X .
 $d_{tr}X_i$ is the quantity involved in an (internal) transformation flow, at a thermodynamic state, i , before or after the transformation.
- (7) IUPAC examples: (electrical) conductance, $G=1/R$ [$C^2\cdot W^{-1} = S$], resistance, R .
The conductance is the ratio of a flow, I_{tr} , and a scalar force, F_{tr} . The symbol R_{tr}^{-1} is chosen to avoid confusion of G for Gibbs energy and $G=R^{-1}$ for conductance.

TABLE A2. Specific quantities, volume-specific for scalar and area-specific for vector quantities, and the corresponding ergodynamic forces. Compare Table A1.

name	symbol	relation	SI unit	notes
1' specific power scalar vector	$P_{V, \text{tr}}$	$= d_{\text{tr}} G_V / dt$ $= J_{\text{tr}} F_{\text{tr}}$ $= J_{\text{tr}} F_{\text{tr}}$	$\text{W} \cdot \text{m}^{-3}$	1
2' specific Gibbs energy scalar vector	$d_{\text{tr}} G_V$	$= d_{\text{tr}} G V^{-1}$ $= d_{\text{tr}} Y F_{\text{tr}}$ $= d_{\text{tr}} Y F_{\text{tr}}$	$\text{J} \cdot \text{m}^{-3} = \text{kg} \cdot \text{m}^{-1} \cdot \text{s}^{-2}$	1
3 ergodynamic force, scalar	F_{tr}	$= \partial_{\text{tr}} G_V / \partial_{\text{tr}} Y$	$\text{J} \cdot \text{x}^{-1}$	2
3' vector	F_{tr}	$= \partial_{\text{tr}}^2 G_V / (\partial_{\text{tr}} Y \partial s)$	$\text{N} \cdot \text{x}^{-1}$	3
4' flux, scalar	J_{tr}	$= I_{\text{tr}} V^{-1}$	$\text{x} \cdot \text{s}^{-1} \cdot \text{m}^{-3}$	4
vector	J_{tr}	$= I_{\text{tr}} A^{-1} = I_{\text{tr}} ds/dV$	$\text{x} \cdot \text{s}^{-1} \cdot \text{m}^{-2}$	5
5' specific advancement, scalar	$d_{\text{tr}} Y$	$= d_{\text{tr}} \xi V^{-1}$	$\text{x} \cdot \text{m}^{-3}$	
vector	$d_{\text{tr}} Y$	$= d_{\text{tr}} \xi A^{-1} = d_{\text{tr}} \xi ds/dV$	$\text{x} \cdot \text{m}^{-2}$	
6' transformant concentration, transformant density	$X_{V, \text{tr}}$	$= X_{\text{tr}} V^{-1}$	$\text{x} \cdot \text{m}^{-3}$	6
surface transformant density	$X_{A, \text{tr}}$	$= X_{\text{tr}} A^{-1}$	$\text{x} \cdot \text{m}^{-2}$	7
7' conductivity, scalar	L_{tr}	$= -J_{\text{tr}} F_{\text{tr}}^{-1}$	$\text{x}^2 \cdot \text{W}^{-1} \cdot \text{m}^{-3}$	
vector	L_{tr}	$= -J_{\text{tr}} F_{\text{tr}}^{-1}$	$\text{x}^2 \cdot \text{W}^{-1} \cdot \text{m}^{-1}$	8

(1) IUPAC: The adjective *specific* before the name of an extensive quantity is often used to mean *divided by mass*.

Specific quantities are obtained by division of the extensive quantity by the size of the system. *Volume-specific* quantities are indicated by subscript V , e.g. $G_V = G V^{-1}$ [$\text{J} \cdot \text{m}^{-3}$]. *Mass-specific* quantities are indicated by subscript m , e.g. $G_m = G m^{-1}$ [$\text{J} \cdot \text{g}^{-1}$]. Specific molar quantities are indicated by subscript n or n_B , e.g. $G_{n_B} = G n_B^{-1}$ [$\text{J} \cdot \text{mol}^{-1}$]. *Specific molar* quantities and intensive molar quantities must be distinguished. A specific molar quantity is related to the amount of substance B contained in the system, n_B (e.g. amount of protein as an expression of 'biomass'). In contrast, an intensive molar quantity is related to the amount of substance (e.g. $d_{\text{tr}} n_B$) involved in a chemical transformation (e.g. $\Delta_{\text{tr}} G_B = F_{B, \text{tr}} = \partial G / \partial_{\text{tr}} n_B v_B$).

(2) Ergodynamic scalar forces are potential differences (Table A1).

(3) Ergodynamic vector forces are gradients of potential, for example $\text{grad} V$, $\text{grad} \mu$.

(4) IUPAC example: rate of reaction, v .

(5) IUPAC examples: electric current density, j , J ; flux (of a quantity X), $J_X = dX/dt A^{-1}$; heat flux, J_q .

(6) IUPAC examples: (electric) charge density, ρ ; (chemical) amount concentration, c_B .

(7) IUPAC example: (electric) surface charge density, σ .

(8) IUPAC example: (electric, electrochemical) conductivity, κ , γ , σ .

TABLE A3. Electric (electrochemical) transformations, el, involving electrical potential differences.

name	symbol	relation	SI unit	notes
1 electric power	P_{el}	$= I_{el} F_{el}$	$W=A \cdot V$	
1' specific electric power	$P_{V,el}$	$= J_{el} F_{el}$	$W \cdot m^{-3}$	
2 electric energy	$d_{el}G$	$= d_{el}\xi F_{el}$	$J=C \cdot V$	1
3 electric force	F_{el}	$= \partial G / \partial_{el}\xi = V_2 - V_1$	V	2
3' vector electric force	\mathbf{F}_{el}	$= \text{grad } V_{el}$	$V \cdot m^{-1}$	3
electric potential	V_{el}	$= \partial G / \partial Q_{el}$	V	4
4 electric flow	I_{el}	$= d_{tr}\xi_{el}/dt = -R_{el}^{-1} F_{el}$	$A=C \cdot s^{-1}$	5
4' scalar electric flux	J_{el}	$= I_{el} V^{-1}$	$A \cdot m^{-3}$	
4' vector electric flux	\mathbf{J}_{el}	$= I_{el} A^{-1} = -L_{el} F_{el}$	$A \cdot m^{-2}$	6
5 electric advancement	$d_{el}\xi$	$= d_{el}Q_i v_i^{-1}$	C	7
5' specific electric advancement	$d_{el}Y$	$= d_{el}\xi V^{-1}$	$C \cdot m^{-3}$	
6 electric charge	Q_{el}		C	8
6' electric charge density	$Q_{V,el}$	$= Q_{el} V^{-1}$	$C \cdot m^{-3}$	9
6' surface charge density	$Q_{A,el}$	$= Q_{el} A^{-1}$	$C \cdot m^{-2}$	10
7 electric conductance	R_{el}^{-1}	$= -I_{el} F_{el}^{-1}$	$S=A \cdot V^{-1}=C^2 \cdot W^{-1}$	11
7' electric conductivity	L_{el}	$= -\mathbf{J}_{el} F_{el}^{-1}$	$S \cdot m^{-1}$	12

- (1) The partial electric energy change is negative for a spontaneous (exergonic) electric process.
- (2) IUPAC: electric potential difference, $U = V_2 - V_1$, ΔV , $\Delta\phi$; electromotive force, $E = \int (F/Q) ds$; (electrochemistry) electromotive force of the cell, potential (*difference*) of the cell reaction, $E = E^0 - (RT/nF) \times \sum v_i \ln a_i$; $E^0 = -\Delta_r G^0/nF$.
 F_{el} is negative for spontaneous (exergonic) electric or electrochemical advancement.
- (3) IUPAC: electric field strength, $\mathbf{E} = \mathbf{F}/Q = -\text{grad}V$.
Note that ergodynamic vector forces are gradients of potentials, therefore the sign is reversed relative to \mathbf{E} .
- (4) IUPAC: electric potential, $V = dW/dQ$, ϕ .
- (5) IUPAC: electric current, $I = dQ/dt$.
Ohm's law is a flow/force relation, the (scalar) force being the electric potential difference.
- (6) IUPAC: electric current density, j , \mathbf{J} ; $I = \int \mathbf{J} \cdot d\mathbf{A}$; $d\mathbf{A}$ is a vector element of area.
- (7) The quantity involved in the flow (negative or positive charge) must be defined.
- (8) IUPAC: quantity of electricity, electric charge, Q ; (electrochemical) elementary charge, proton charge, e .
- (9) IUPAC: charge density, $\rho = Q/V$.
- (10) IUPAC: surface charge density, $\sigma = Q/A$.
- (11) IUPAC: (electric) conductance, $G = 1/R = I/U$.
- (12) IUPAC: (electric) conductivity, κ , γ , σ ; (electrochemical) conductivity, $\kappa = j/E$.

TABLE A4. Thermal transformations, th, involving temperature differences.

<i>name</i>	<i>symbol</i>	<i>relation</i>	<i>SI unit</i>	<i>notes</i>
1 thermal power	P_{th}	$= I_{th} F_{th}$	W	
2 thermal energy (exergy)	$d_{th}G$	$= d_{th}\xi F_{th}$	J	
3 thermal force	F_{th}	$= (T_2-T_1)/T_1$	1	1
4 heat flow	I_{th}, I_Q	$= d_{th}\xi/dt$	W	2
4' heat flux	J_{th}	$= I_{th} A^{-1}$	W·m ⁻²	3
5 thermal advancement	$d_{th}\xi$	$= d_{th}Q_i v_i^{-1}$	J	
6 heat	Q		J	
7 thermal conductance	R_{th}^{-1}	$= -I_Q [(T_2-T_1)/T_1]^{-1}$	W	4

- (1) $-F_{th}$ is explained as 'maximum machine efficiency' or 'Carnot efficiency' in classical thermodynamics. Since the flow-force product is defined to yield power [W], heat flow [J·s⁻¹ = W] is conjugated to a dimensionless force. Force is the partial derivative of potential energy per advancement of heat conduction (see also Table A6).
- (2) IUPAC: heat flow rate, $\phi = dq/dt$.
- (3) IUPAC: heat flux, $J_q = \phi/A$.
- (4) IUPAC: thermal conductance, $G = \phi/\Delta T$ [W·K⁻¹].
 G is defined differently from R_{th}^{-1} .

TABLE A5. Chemical reactions, r, involving differences of stoichiometric potential.

<i>name</i>	<i>symbol</i>	<i>relation</i>	<i>SI unit</i>	<i>notes</i>
1 reaction power	P_r	$= I_B F_B$	W	
1' specific reaction power	$P_{V,r}$	$= J_B F_B$	W·m ⁻³	
2 reaction Gibbs energy	$d_r G$	$= d_r \xi_B F_B$	J	
2' specific reaction Gibbs energy	$d_r G_V$	$= d_r G V^{-1}$	J·m ⁻³	
3 Gibbs force	$F_{B,r} (\Delta_r G_B)$	$= (\partial G/\partial_r \xi_B)_{T,p}$ $= \sum_i v_i G_i$	J·mol ⁻¹	1
molar reaction enthalpy	$\Delta_r H_B$	$= (\partial H/\partial_r \xi_B)_{T,p}$ $= \sum_i v_i H_i$	J·mol ⁻¹	2
chemical potential of i , partial molar Gibbs energy	G_i, μ_i	$= (\partial G/\partial n_i)_{T,p,n_{j \neq B}}$	J·mol ⁻¹	
partial molar enthalpy of i	H_i	$= (\partial H/\partial n_i)_{T,p,n_{j \neq B}}$	J·mol ⁻¹	
4 reaction flow	$I_{B,r}, I_B$	$= d_r \xi_B/dt$	mol·s ⁻¹	3
4' reaction flux	$J_{B,r}, J_B$	$= d_r \xi_B/dt V^{-1}$ $= d_r Y_B/dt$	mol·s ⁻¹ ·m ⁻³	4

TABLE A5. (cont.)

name	symbol	relation	SI unit	notes
reaction enthalpy flux	$J_{H,r}$	$= J_{B,r} \Delta_r H_B$	$\text{W}\cdot\text{m}^{-3}$	
scalar heat flux	J_Q	$= I_Q V^{-1}$	$\text{W}\cdot\text{m}^{-3}$	5
5 advancement of reaction	$d_r \xi_B$	$= d_r n_i v_i^{-1}$	mol	6
stoichiometric number of substance i in reaction j	v_{ij}			
5' specific extent of reaction	$d_r Y_B$	$= d_r \xi_B V^{-1}$	$\text{mol}\cdot\text{m}^{-3}$	
6 change of reacting amount of i	$d_r n_i$	$= d_r n_i + d_e n_i$	mol	7
amount of substance	n_i	$= m_i M_i^{-1}$	mol	
6' concentration change of reacting substance	$d_r c_i$	$= d_r n_i V^{-1}$	$\text{mol}\cdot\text{m}^{-3}$	8
amount concentration	c_i	$= n_i V^{-1}$	$\text{mol}\cdot\text{m}^{-3}$	

- (1) IUPAC: reaction Gibbs energy (function), $\Delta_r G$, $\Delta_r G_m$; affinity of reaction, $-A = \Delta_r G$. The subscript B in $F_{B,r}$ and $\Delta_r G_B$ indicates that any given reaction stoichiometry is divided by $|v_B|$ such that a stoichiometric form is obtained with $|v_B| = 1$. In matrix notation, F_{ij} indicates the Gibbs force of the j th reaction involving substance i .
- (2) $\Delta_r H_B$ refers to a stoichiometric form with the modulus of the stoichiometric number of B of unity. It is a process-intensive quantity but not an ergodynamic force. In matrix notation, the symbol $\Delta_r H_{ij}$ is the molar enthalpy of the j th reaction involving substance i .
- (3) IUPAC: rate of conversion, $\dot{\xi} = d\xi/dt$.
- (4) IUPAC: rate of reaction, ν . Note that the symbol ν (velocity) is different from Greek nu, ν , for stoichiometric number. In an earlier edition, the symbol J was recommended by IUPAC for the 'rate of reaction', $d_r \xi/dt$ (ref. 15; corresponding to reaction flow, I_B). In ref. 1 the alternative symbol is dropped and the term changed to 'rate of conversion', whereas 'rate of reaction' is now $d_r \xi/dt V^{-1}$ (corresponding to reaction flux, J_B).
- (5) J_Q is the calorimetrically measured heat flux, equal to $J_{H,r}$ in the absence of work and of any other transformations.
- (6) IUPAC: extent of reaction, advancement, $\Delta \xi = \Delta n_B / \nu_B$. The subscript r indicates that (internal) transformations of i due to a reaction (in general) are considered. $d_r n_i$ is negative if i is a substrate, positive if i is a product. The advancement $d_r \xi_B = d_r n_i v_i^{-1}$ is positive when B is a substrate or a product when the reaction proceeds in the direction as defined by the stoichiometry. The subscript B indicates the stoichiometric form with $|v_B| = 1$.
- (7) dn_i is the change of the amount of i contained in the system. $d_r n_i$ is the internal change. $d_e n_i$ is the external change of the amount of i in an open system. $d_r n_i = -d_e n_i$ holds for a reaction at steady state.
- (8) IUPAC: rate of concentration change (due to chemical reaction), $r_B = dc_B/dt$, ν_B .

TABLE A6. Important ratios related to stoichiometry and efficiency.

name	symbol	relation	SI unit	notes
stoichiometric number of substance i in transformation j	v_{ij}			
of substance i per unit B	$v_{i/B}$	$= v_i v_B ^{-1}$	$\text{mol } i \cdot \text{mol}^{-1} \text{ B}$	1
flux ratio of a quantity <i>out</i> per quantity <i>in</i>	$Y_{\text{out/in}}$	$= J_{\text{out}}/J_{\text{in}}$	variable	2

TABLE A6. (cont.)

name	symbol	relation	SI unit	notes
heat flux/reaction flux ratio	$Y_{Q/B}$	$= J_Q/J_B$	J·mol ⁻¹	3
heat flux/enthalpy flux ratio	$Y_{Q/H}$	$= J_Q/J_H$		4
normalized flux ratio	j	$= Y_{out/in} v ^{-1}$	variable	5
normalized force ratio	f	$= -F_{out}/F_{in} v $		5
efficiency	ϵ	$= -P_{out} P_{in}^{-1} = j f$		6
ergodynamic efficiency		$= -Y_{out/in} F_{out}/F_{in}$		
ergodynamic efficiency of a heat engine	ϵ	$= \eta/\eta_{max}$		7
heat engine efficiency	η	$= -P_{out} I_{th}^{-1} = -d_e W_{out} d_e Q_1^{-1}$		7

- (1) $\nu_{i/B}$ is the stoichiometric number of substance i in a reaction with a stoichiometric number of B of unity. $\nu_{i/B}$ is positive for a product i and negative for a substrate i ,

$$\nu_{i/B} = \nu_i | \nu_B |^{-1}$$

- (2) A molar flux ratio of substance P per S, $Y_{P/S}$, is (1) the flux ratio of two coupled reactions $d_r Y_{P,2}/d_r Y_{S,1}$. If there is an invariant stoichiometry between the changes of P and S, e.g. ATP and glucose, then the flux ratio is constant, reactions r1 and r2 can be written as a single fully coupled reaction r, and $Y_{P/S}$ is the absolute value of a stoichiometric number, $| \nu_{P/S} |$. (2) $Y_{P/S}$ is the flux ratio in a complex stoichiometric reaction, such as the reaction equation for microbial growth on a given substrate S. Then $Y_{P/S}$ is a flux ratio, conventionally known in biotechnology as 'yield'.
- (3) An important example of the heat flux/reaction flux ratio, $Y_{Q/B}$, is the calorimetric/respirometric ratio (CR ratio), where substance B is the oxygen consumed in respiration,

$$\text{CR ratio} = Y_{Q/O_2} = J_Q / J_{O_2}$$

The heat flux/reaction flux ratio, Y_{Q/O_2} , is contrasted with the molar enthalpy of reaction k, $\Delta_k H_{O_2}$. The former is related to a direct calorimetric experiment (usually in a complex reaction system), whereas the latter is thermochemically derived, e.g. calculated from enthalpies of formation.

- (4) The heat/enthalpy flux ratio is important in calorimetric and metabolic enthalpy balance studies. If all reactions and side reactions in a complex reaction system which does not exchange energy in the form of work are accounted for, the heat/enthalpy flux ratio equals 1. If $Y_{Q/H} < 1$, then the thermochemical analysis of energy transformations failed fully to account for the calorimetrically measured heat flux. $Y_{Q/H} > 1$ indicates that some endothermic transformations remained undetected. $Y_{Q/H}$ can be interpreted as a normalized CR ratio for the example in note 3.
- (5) In two coupled reactions, ν is the stoichiometric number in a reference reaction, e.g. of product P (such as ATP) with ν_S of unity (e.g. glucose; O_2). Then $\nu_{P/S}$ is the stoichiometry, with reference to which the experimentally observed flux ratio $Y_{P/S}$ is normalized. The reference stoichiometry, ν , may also be symbolized as P:S (for example ATP:Glucose; ATP: O_2).
- (6) The power efficiency or ergodynamic efficiency, ϵ , equals the -output/input Gibbs energy ratio.
- (7) The thermodynamic or heat engine efficiency, η , is the ratio of work done on the surroundings, $d_e W_{out}$, to the external heat input at the higher temperature, $d_e Q_1$. At steady state this equals the ratio of internal power output, P_{out} , divided by the internal heat flow, I_{th} . The maximum thermodynamic efficiency of a Carnot engine is the Carnot efficiency, $\eta_{max} = -F_{th}$ (compare Table A4),

$$\eta_{max} = \frac{T_1 - T_e}{T_1}$$

where T_1 and T_e are the high and low temperatures of the heat source (1) and external heat sink (e), respectively, between which the Carnot engine operates.