

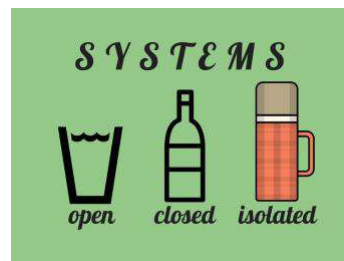
System Definitions

- System: the part of the universe being studied (big –Earth, or small –one atom)
- Surroundings: the part of the universe outside the system
- We look at transfer of energy/matter between system and surroundings.

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System Definitions

- Open system: free exchange of matter and energy with surroundings
- Closed system: free exchange of energy only
- Isolated system: no interaction with surroundings



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Conservation of Energy

- In interactions between system and surroundings, **energy cannot be created or destroyed, only transferred**
- Total energy (Q) will remain constant
$$Q_{\text{system}} + Q_{\text{surroundings}} = 0$$
- Basically: all energy that leaves the system goes into the surroundings, and vice versa

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Conservation of Energy

- In other words, heat lost by one object is gained by the other

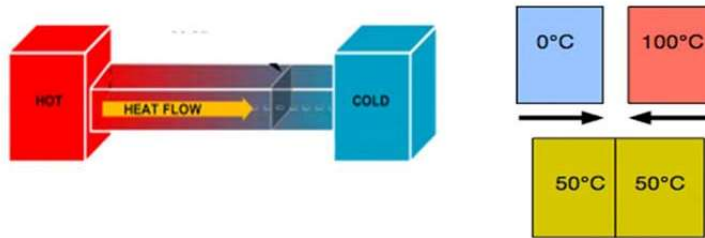
$$Q_1 = -Q_2$$

- The total heat Q of the two objects are 0

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Thermal Equilibrium

- Is obtained when touching objects within a system reach the same temperature.



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0.50 kg of water at 50.0°C is mixed with 0.20kg of water at 20.0° C. What is the final temperature? Assume there is no heat loss to the surroundings.

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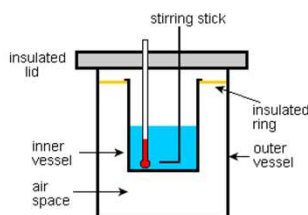
Calorimetry

- The science of determining changes in energy of a system by measuring heat exchanged with surroundings
- Calorimeter: well-insulated vessel (as close to an isolated closed system as possible)

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Calorimetry

- Inside, a heated or cooled object (or chemical reaction) – the system – is submerged in water – the surroundings. The temperature change of the water can be used to determine the heat released by the system.
- The calorimeter needs to prevent heat loss to the surroundings so the temperature change within it can be accurately recorded.



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A calorimeter contains 0.50 kg of water at 15°C. A 0.040 kg block of Zn at 115°C is placed in the water. What is the final temperature of the system?